EXPERIMENTAL EDITION

SCHOOL CHEMISTRY

CLASSES 6 & 7

TEXT

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1968

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RESEARCH AND TRAINING

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SCHOOL CHEMISTRY

CLASSES 6 & 7

TEXT

Prepared by

NCERT Chemistry Study Groups

Under a grant from

THE NATIONAL COUNCIL OF EDUCATIONAL
RESEARCH AND TRAINING
NEW DELHI

PREFACE

The Education Commission recommended that teaching of science should commence at class 5 with physics and biology as separate disciplines and at class 6 with chemistry. This recommendation has been accepted by the Ministry of Education, Government of India and in pursuance of this, the National Council of Educational Research and Training has set up Study Groups to draw up a curriculum and develop teaching materials in various disciplines.

The Chemistry Study Groups, after long deliberation, considered it necessary that a new approach should be made to the study of chemistry at the school level. Essentially this approach is based upon pupil experimentation followed by teacher's demonstration and a discussion leading to an understanding of the principles of chemistry. Accordingly the laboratory manual and text have been prepared for classes 6 and 7. The pupil is expected to carry out the experiments described in the laboratory manual either individually or in groups as indicated. In this approach, the pupil, apart from learning about the behaviour of substances, has also the thrill of finding something new all by himself. The teacher's demonstrations supplement the experimental work done by the pupils and lead to further discussion of the basic principles of chemistry.

For conducting these experiments it is not absolutely necessary to have a separate laboratory, they can be carried out in a class room provided with suitable flat desks. The equipment needed is also of a simple and inexpensive nature.

The project was started in September 1966 and after several meetings of the Study Groups, a broad outline of the curriculum was drafted and later, the writing of the Laboratory Manual, Text and Teacher's Guide was undertaken. Some of the material prepared was tried out in a few schools, outside the normal curriculum to find out the response from the pupils. The response has been encouraging, especially because of the experimental work involved.

The Directors and members of the Study Groups are conscious of the shortcomings and limitations of the material, which was prepared by assuming a certain level of knowledge on the part of the teacher. The practical difficulties in implementing the course will be obvious only after trial for a full academic year. For this reason the present material is being brought out as an experimental edition which will undergo revision in the light of the experience gained by trying it out in a number of schools. Helpful comments and suggestions for improvement of the Laboratory Manual as well as the Text will be gratefully received.

I will be failing in my duty, if I do not acknowledge the valuable help given by the Directors and members of the staff of the Study Groups as well as the staff of the Department of Science Education, National Institute of Education, New Delhi, in the preparation of this experimental edition. But for the financial support given by NCERT and encouragement by Dr. D. S. Kothari, Chairman, University Grants Commission and Mr. L. S. Chandrakant, the then Joint Director of the NCERT, the project would not have taken shape The Directors and members of the Study Groups gratefully acknowledge their help and encouragement.

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UNIT 1

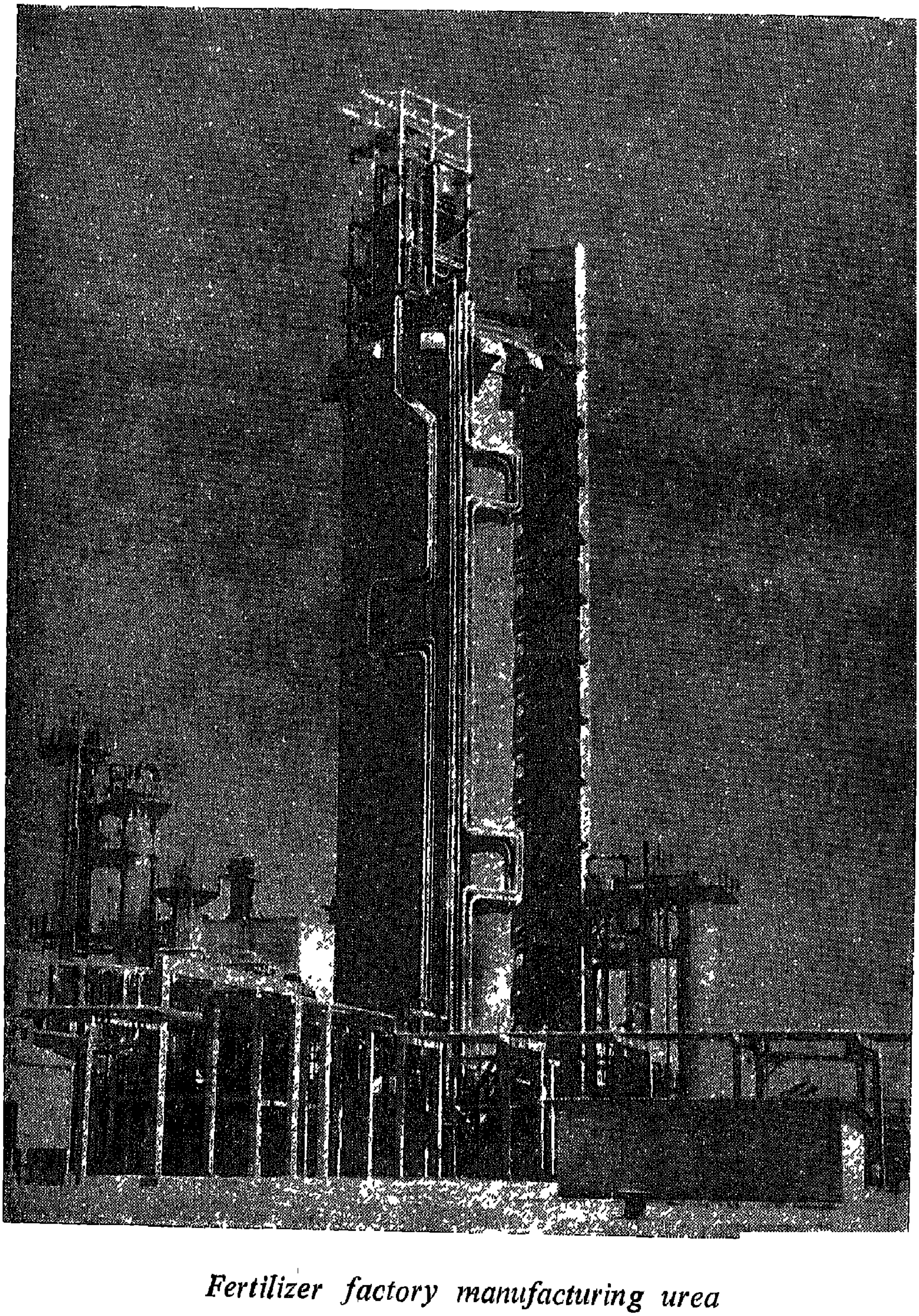
Chemistry and the Material World

In nature we come across a variety of substances. This unit deals with the properties and composition of different substances and the changes they undergo. Chemistry helps us in a number of ways in our daily life.

1.1 What is chemistry?

The world around us is composed of different substances. Our earth is surrounded by a vast blanket of air and a major portion of its surface is covered with water. Sand, coal, iron, copper, salt, sugar and kerosene are some of the common substances we use in daily life. These substances have different properties. For example, coal is black, sulphur yellow, copper reddish-brown and salt and sugar white. Water and kerosene flow freely. Salt and sugar disappear when placed in water imparting to it a characteristic taste. We can burn kerosene but not water. Iron and stone are hard and heavy while charcoal is light and soft. Thus in nature there is a large number of substances with different properties.

Substances undergo changes. For example, coal burns, milk sours, iron rusts and water evaporates. Why does coal burn? Why does iron rust? Many such questions arise in our daily life. Answers to these can be obtained by a *systematic study of materials—finding out what they are made up of and the changes they undergo*. This is what **chemistry** deals with.



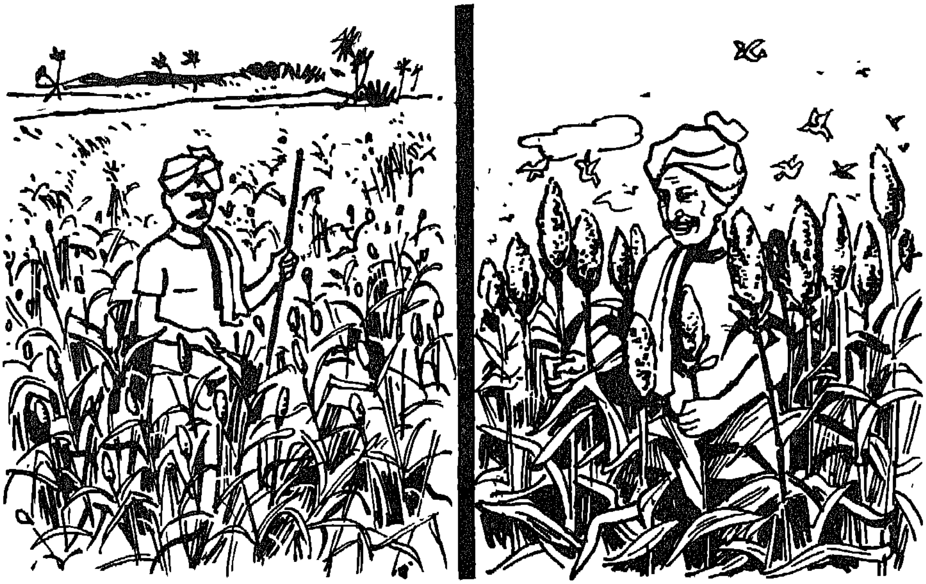
Fertilizer factory manufacturing urea

1.2 Importance of chemistry in daily life

The paper, the ink, and the type metal used to prepare the book you read, the pen you write with, the clothes you wear, the soap you use, the drugs which a doctor prescribes, the cement that is used in building houses, the sugar that is used as a food material are all produced with the knowledge of chemistry.

Let us see how chemistry helps us in different spheres of life.

Food is essential for all living beings. Plants make food for us. Fertilizers like ammonium sulphate, urea and superphosphate help in the proper growth of plants and thereby give better yield of crops (Fig. 1.1).



Crop without fertilizer *with fertilizer*

Figure 1.1 *Effect of fertilizers on plant growth*

Plants, during their growth, may be attacked by harmful insects and fungi. They have to be protected from these, otherwise the yield of the crops will be poor. Crops are protected by spraying chemicals like D.D.T., benzene hexachloride, sulphur, etc. (Fig. 1.2).

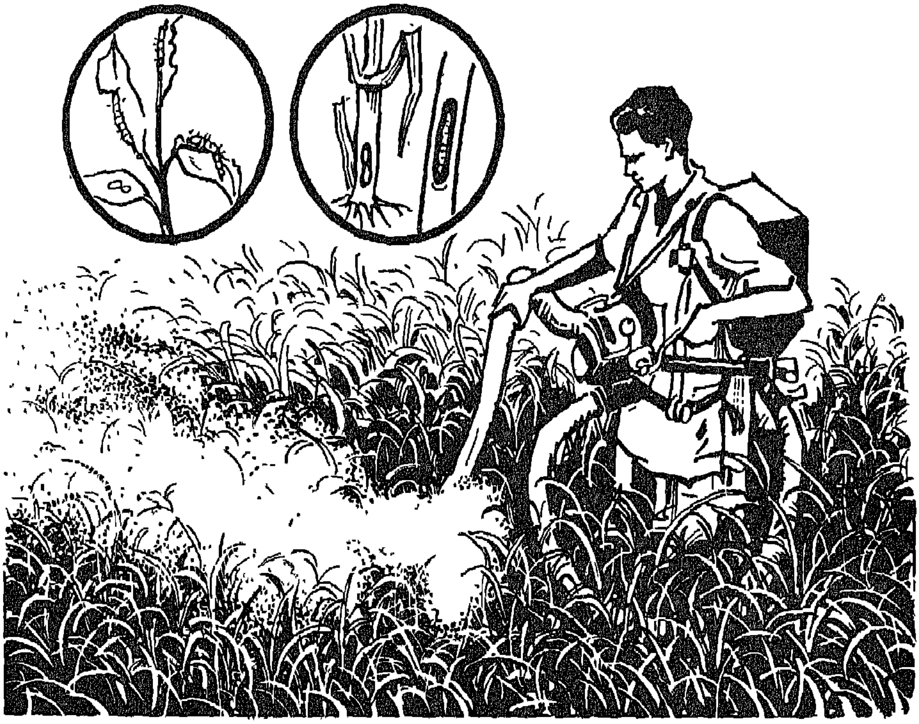


Figure 1.2 *Protection of plants by dusting pesticides*

The food produced has to be used over a long period. Some of it will be used in the off season. Sometimes it has to be transported to other places. To keep food for a long time we have to preserve it. For example, salt is used for preserving pickles and fishes. Sugar is used for preserving fruits in the form of jams and jellies Foods are also preserved by dehydration and canning (Fig. 1.3).

Vitamins are essential for good health. The lack of vitamins causes diseases such as night blindness, beriberi, scurvy and rickets. Most vitamins are found in our daily food (Fig. 1.4b). Sometimes the food we take may not have the required amounts of vitamins. To make up for this deficiency our food has to be supplemented by various vitamins. These have been prepared with the knowledge of chemistry.

Figure 1.3 *Preserved foodstuffs*

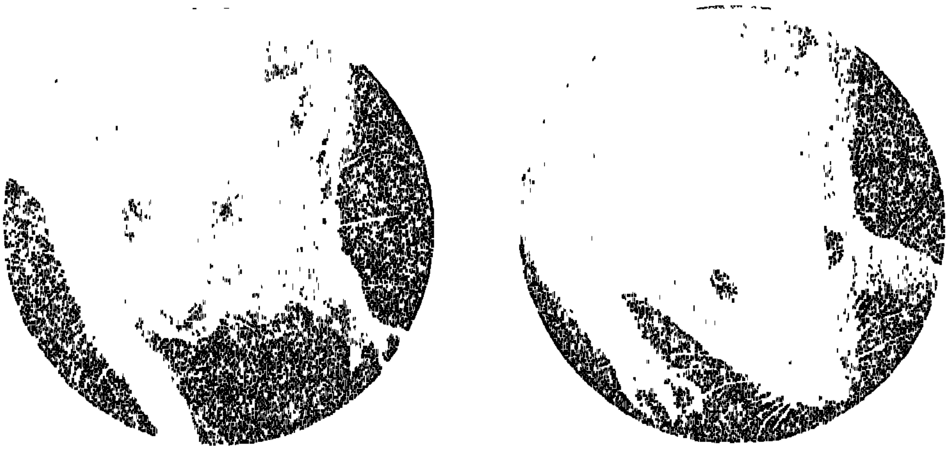


Figure 1.4 a

Mouse suffering from vitamin A deficiency

After being given vitamin A

Figure 1.4b *Some foodstuffs rich in vitamins*

We owe a great deal to chemistry for making our lives healthier. Medicines like aspirin, sulphadrugs, penicillin and other antibiotics are used to fight diseases. Disinfectants like bleaching powder, potassium permanganate and phenyle are used to prevent the spread of diseases.

Cotton, woollen and silk clothes prepared from natural fibres have been in use from times immemorial. Knowledge of chemistry has helped us in preparing new fibres like rayon, nylon or terylene to meet the ever increasing demand for clothing. To impart different colours and shades to clothing, dyes have also been synthesised and are fast replacing the natural ones.

The production of materials such as glass, cement, mortar, asbestos, tiles and porcelain used in construction of buildings involves a knowledge of chemistry. Iron and steel are used to strengthen building structures (Fig. 1.5).

Figure 1.5 *Materials used in building construction*

Fuels like coal, diesel and petrol help us in developing quicker transport (Fig. 1.7).

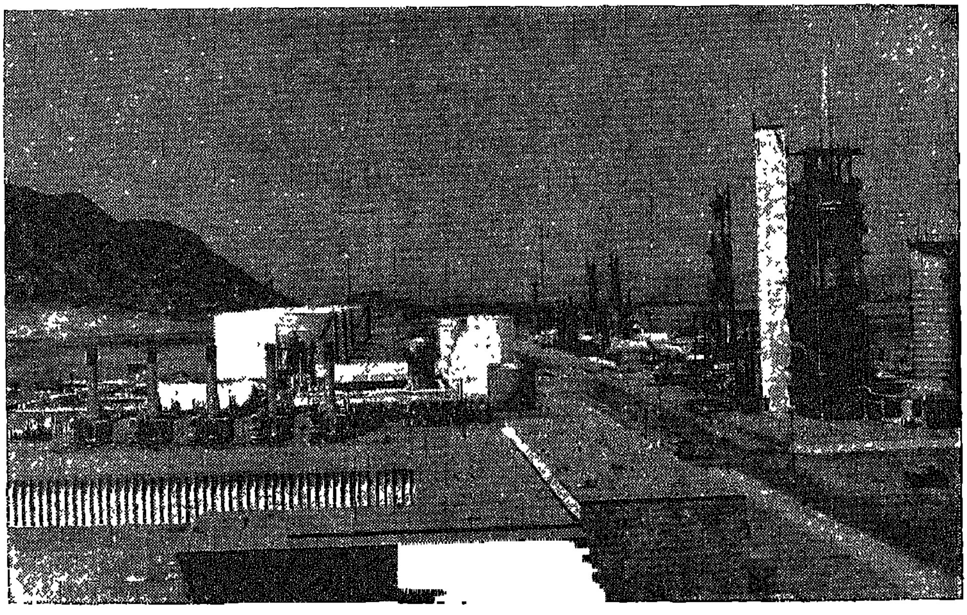


Figure 1.6 *Petroleum refinery*

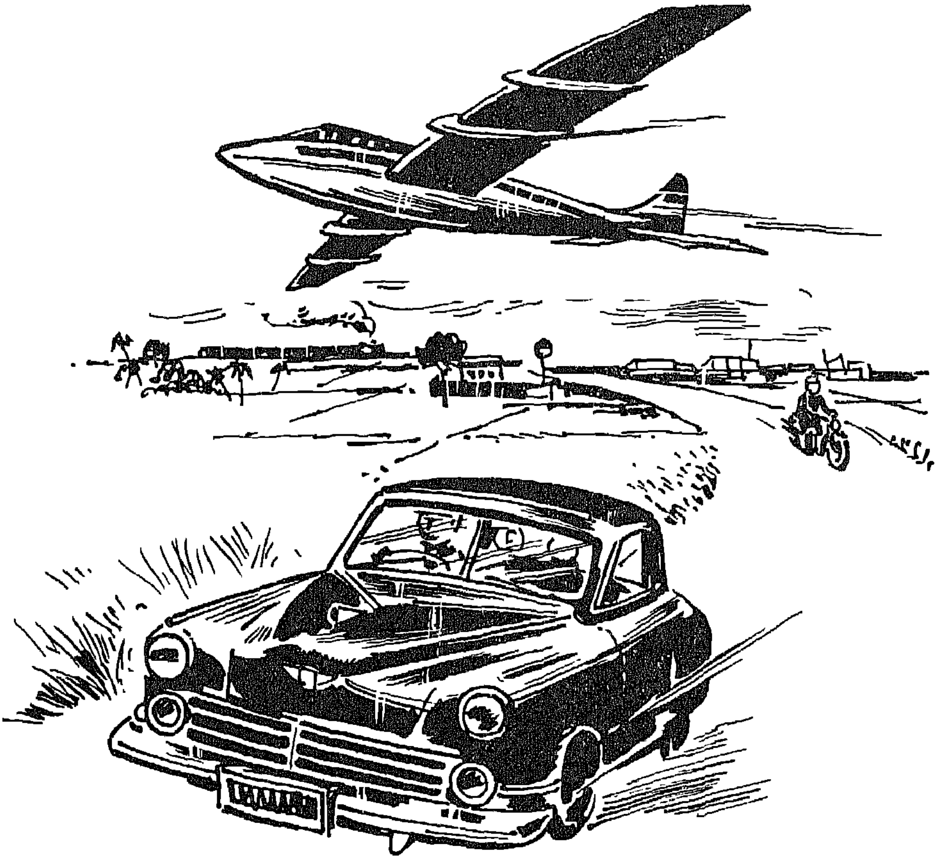


Figure 1.7 *Modern transport using liquid fuels*

Thus chemistry helps in

1. providing more and better food, clothing and housing,
2. developing substitutes for essential materials such as rubber, textiles and natural drugs,
3. making substances that are entirely new.

The study of chemistry gives a fund of information about materials and processes useful in everyday life. A knowledge of chemistry is essential in many professions. For instance it is vital for a doctor or a pharmacist. A farmer should know what type of fertilizers and pesticides are needed to improve the yield of crops. For technologists engaged in the extraction of metals from their ores, refining of petroleum, manufacture of drugs and various commodities of daily use, an adequate knowledge of chemistry is necessary (Fig.1·8).

Figure 1.8 *Chemistry in the service of man*

1.3 Matter and the changes it undergoes

Matter

Air, water, a book, a chair and an iron bar are a few examples of matter. These occupy space and possess mass. *Anything that occupies space and possesses mass is called* **matter**. Matter can exist in one of the three states—solid, liquid or gaseous. We all know that solids and liquids have mass (lift a bucket of sand or water and feel how heavy it is). What about air ? Does it also have any mass ? Let us find out.

Two balloons filled with air are suspended, one on each end of a meter scale and balanced. One of them is now punctured to allow the air to escape. The end carrying the punctured balloon is found to go up.

What made it go up ? The escape of the air which filled the balloon This indicates that the other balloon still containing air is heavier than the punctured one. From this we conclude that air possesses mass. Similarly all other gases possess mass.

Properties of matter

We observe that solids like brick, wood and book retain their shape and volume wherever we place them. However, a measured volume of a liquid such as water when poured into vessels of different shapes like a beaker, a conical flask and a flat-bottomed flask, assumes the shape of its container. This shows that water has no fixed shape even though its volume is the same. This is true of all liquids.

Sodium thiosulphate reacts with nitric acid giving a brown gas, nitrogen dioxide. A few crystals are placed on three watch glasses. The crystals are treated with a few drops of concentrated nitric acid. Three glass vessels of different sizes and shapes are inverted over them. After some time brown vapours of the gas are found to fill each vessel completely. This shows that the gas spreads throughout the available space and assumes the shapes of the vessels. Thus gases do not have a definite volume or a definite shape. Similarly, air pumped into balloons of different shapes assumes the shapes of the balloons.

Objects like the wooden block *which retain a definite shape and volume are called* **solids.** *Substances* like water *which retain their* *volume but change their shapes are called* **liquids.** *Substances* like air or the brown gas *which occupy any volume and assume any shape are called* **gases.**

Materials we see around us exist in one of these states.

Can a substance exist in all the three states ?

A few pieces of ice (solid) when taken in a large test tube and warmed melt to form water (liquid). Further heating causes water to boil and give out steam (gas). On cooling, steam condenses to water which, on further cooling in a freezing mixture, gives ice.

$$\underset{\text{(solid)}}{\text{Ice}} \quad \underset{\text{cool}}{\overset{\text{heat}}{\rightleftarrows}} \quad \underset{\text{(liquid)}}{\text{Water}} \quad \underset{\text{cool}}{\overset{\text{heat}}{\rightleftarrows}} \quad \underset{\text{(gas)}}{\text{Steam}}$$

These changes show that water exists in all the three states of matter. Many other substances behave similarly. For example, tin metal is a white shining solid On heating, it will melt and eventually turn into vapour. Even an iron bar can be melted and vapourised at high temperatures.

Not all substances exist in all the three states. For example, it is observed that ammonium chloride does not melt on heating but vapourises directly. The vapours are condensed on cooler parts of the test tube as a solid deposit. Such a change is called **sublimation.** Camphor and iodine also sublime when heated.

Types of matter

As mentioned earlier, air, sea water and sugar are common examples of matter. These have different properties. They may contain one or more substances either intimately mixed or in a combined form. Matter is classified into different types on the basis of its properties and chemical composition

Mixtures

Iron and sulphur have different properties. Iron is attracted by a magnet whereas sulphur is not. Sulphur is soluble in carbon disulphide but iron is not. When iron filings and sulphur are mixed together, a sample with the following characteristics is obtained :

a. The particles of iron and sulphur are seen separately.

b. On moving a magnet through this sample only iron particles stick to the magnet leaving sulphur behind.

c. On shaking a small amount of the sample with carbon disulphide, sulphur alone dissolves leaving iron. Evaporation of the clear solution thus obtained gives sulphur.

These observations show that:

i. Iron and sulphur retain their properties even when mixed together.

ii. Iron and sulphur can be separated by simple methods.

Samples containing different amounts of iron and sulphur have properties characteristic of both. Such samples are called **mixtures.** Concrete and soil are some of the common examples of mixtures.

Is air a mixture?

The following experiment will help us know whether air is a mixture. A small piece of yellow phosphorus is taken in a small porcelain dish. The dish is floated on water in a trough. A stoppered bell-jar is placed over it. Above the water level five equal divisions are made on the bell-jar (Fig. 1.9). The stopper is removed and the phosphorus is kindled by a hot iron wire and the stopper is quickly replaced. Phosphorus burns giving off dense white fumes which dissolve in water. It is observed that water level in the bell-jar rises by about one-fifth of the total volume of air

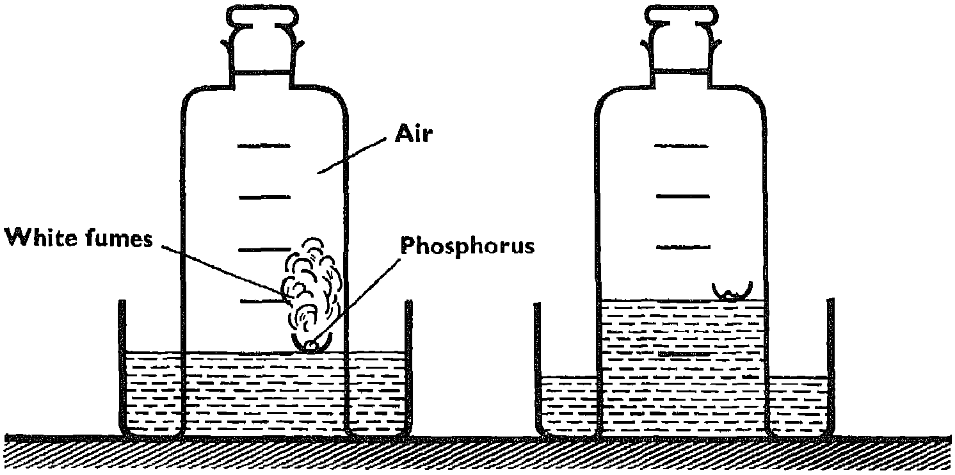


Figure 1.9 *Air is a mixture.*

Some phosphorus is left unburnt. The remaining air puts out a burning splinter. This is the inactive part of air which is mostly nitrogen. It constitutes about four-fifths of the total volume of air. The active part of air used up in the burning of phosphorus forms about one-fifth of the total volume of air. It is oxygen. Thus air contains mainly two gases — nitrogen and oxygen It contains a small proportion of other gases which will be studied later. These gases can be separated from one another. Therefore air is a mixture.

Solutions

Let us now consider another type of mixture. When copper sulphate is added to water it dissolves imparting a blue colour to the mixture. No solid substance is seen. In course of time the colour of the mixture becomes uniform throughout (homogeneous). Such *a homogeneous mixture of two or more substances is called a* **solution.**

When a few drops of sea or well water are evaporated in a watch glass, a residue is left behind. Such water is therefore not a single substance, but a uniform mixture. So it is a solution. Tincture of iodine is also a homogeneous mixture and therefore a solution.

✓

Compounds

Let us see what happens when two substances combine together. Sodium is a bluish white shining solid. Chlorine is a greenish yellow gas. When a piece of sodium is warmed and introduced into a jar of chlorine, a white solid is formed. This white solid has properties different from those of sodium and chlorine. It is a compound formed from sodium and chlorine and is known as sodium chloride (common salt).

On passing a direct electric current through water (Fig. 1.10) containing a little sodium chloride water is decomposed into two gases in the ratio of 2 : 1 by volume. When a burning splinter is introduced into the gas collected in greater volume, the gas burns with a 'pop' sound. This gas is hydrogen. The other gas which rekindles a glowing splinter is oxygen. Water, therefore, is not a simple substance. It is made up of hydrogen and oxygen in the fixed ratio of 2 : 1 by volume. The properties of hydrogen and oxygen are different from those of water. These cannot be separated from water by simple methods. So water is a **compound**.

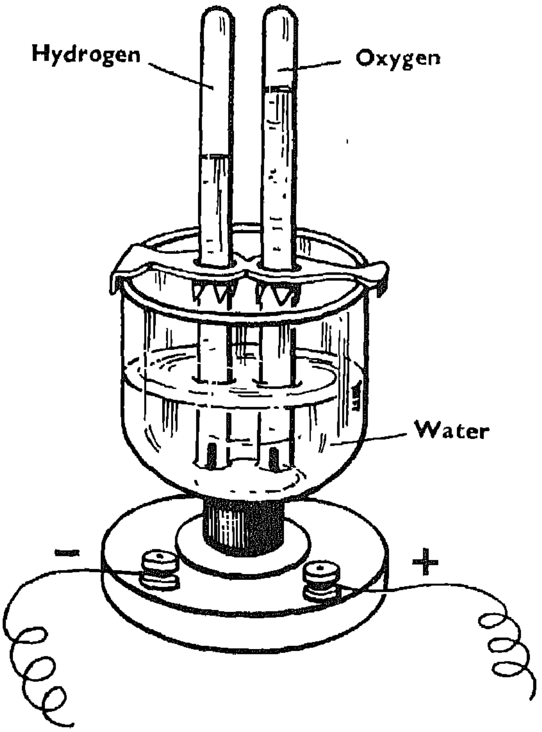


Figure 1.10 *Electrolysis of water.*

Elements

It has been observed that water can be decomposed into two substances — hydrogen and oxygen. These cannot be further decomposed into simpler substances. Such substances are called **elements.** Some of the elements are listed below:

Some Common Elements

Aluminium
Calcium
Carbon
Chlorine
Copper
Gold
Hydrogen
Iodine
Iron
Lead
Magnesium
Mercury
Nitrogen
Oxygen
Phosphorus
Potassium
Silicon
Silver
Sodium
Sulphur
Tin
Zinc

The total number of elements now known is 103. Only a few of these are abundant in nature. The relative abundance of these elements in the earth's crust is shown in Fig. 1.11.

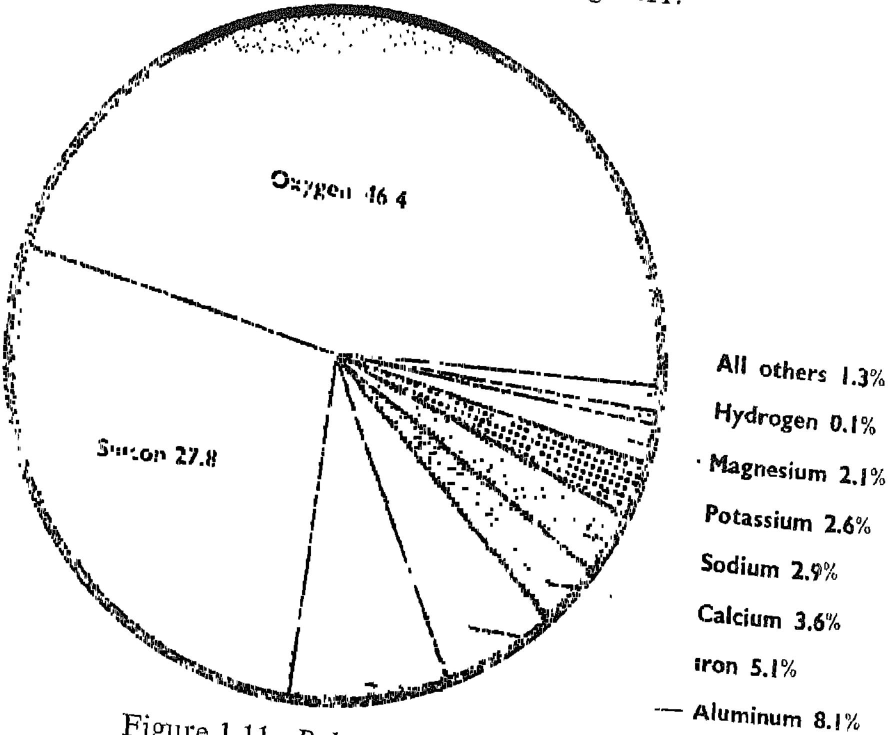


Figure 1.11 *Relative abundance of elements.*

Almost one half of the earth consists of oxygen. Silicon (present in sand and rocks) comprises more than one-fourth. In fact, only nine elements make up a total of over ninetyeight per cent of the earth's crust.

Change of state of matter

When ice is warmed it melts into water, which on cooling gives back ice. Ice and water are therefore the same substance; the difference is that ice is a solid whereas water is a liquid. Melting of ice into water involves only a change of state. No new substance is formed in this process. Such a change is called a physical change.

A platinum or nichrome wire glows on heating. On removing the wire from the flame it regains its original appearance. No new substance is formed. In the sublimation of ammonium chloride (page 11) also, no new substance is formed. These too are examples of physical changes.

When a piece of magnesium ribbon is heated in air, it burns with a dazzling light forming a white ash. This substance has properties different from those of magnesium. It is called magnesium oxide. Magnesium on burning combines with oxygen and forms a new substance, magnesium oxide. Such a change is called a chemical change. Electrolysis of water is also a chemical change.

We come across many chemical changes in our daily life. When we burn charcoal heat is liberated and carbon dioxide is formed. Carbon dioxide is different from charcoal. Hence the burning of charcoal is a chemical change. During rusting, iron is converted into iron oxide. Rusting of iron is also a chemical change.

Conservation of mass

Lead nitrate solution is taken in a conical flask. A test tube containing potassium iodide solution is suspended in the conical flask (Fig. 1.12). Without allowing the solutions to mix, the flask with its contents is weighed. The solutions are then mixed by tilting the flask. A yellow substance is formed. This shows that a chemical change has taken place. On weighing again it is found that there is no change in the weight of the flask and its contents. Thus *during a chemical change there is no change in mass.*

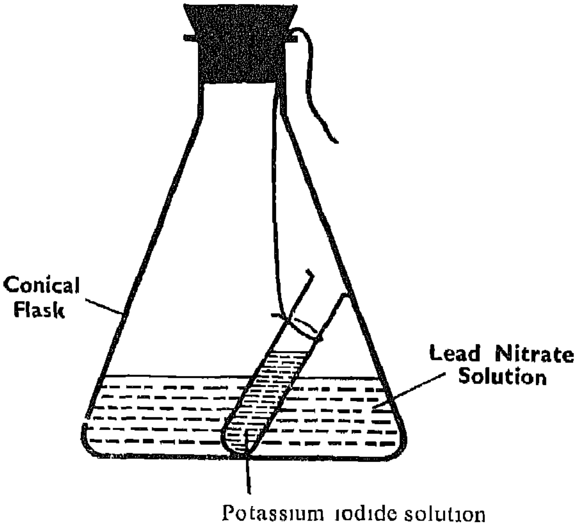


Figure 1.12 *Conservation of mass*

Questions

1. How is chemistry helpful in food production?
2. How does chemistry help in modern transport ?
3. Match the words in columns A and B

A	B
Fertilizer	Rayon
Drug	D. D. T.
Pesticide	Soap
Washing agent	Penicillin
Food preservative	Salt
	Ammonium sulphate

4. Fill up the blanks
   a. Anything that occupies .... and possesses . . . . is described as matter.
   b. Compounds are formed as a result of. ....change
   c. Ammonium chloride ....on heating.
5. Classify the following substances into solids, liquids and gases.
   a. Iron nail
   b. Water
   c. Piece of chalk
   d. Petrol
   e. Oxygen
6. Classify the following materials into mixtures, solutions, compounds and elements.
   a. Sugar
   b. Lime
   c. Soda water
   e. Iron
   f. Gold
   g. Concrete
7. How is water converted into ice ?
8. What is the difference between a compound and an element?
9. What is the most abundant element in the earth's crust ?
10. What are the characteristics of a mixture?
11. On heating, butter is converted into ghee. Is this a physical change or a chemical change ?
12. State whether the following statements are TRUE or FALSE.
    a. Water is a mixture of hydrogen and oxygen.
    b. Air is a compound of nitrogen and oxygen as the proportions of these in air are nearly constant.
    c. Mass changes accompany chemical changes.
    d. Burning of a match stick is a physical change.
    e. Sublimation is a physical change.
    f. Air is matter.
    g. Sea water is a mixture.

UNIT 2

Structure of Matter

Substances behave differently because of differences in their structure. In this unit you will learn that matter is composed of particles called molecules which are built from atoms. Atoms consist of smaller particles called electrons, protons and neutrons.

2.1 Behaviour of gases

We learnt in unit 1 (page 11) that matter exists in three states—solid, liquid and gaseous. A solid has definite volume and shape, whereas a liquid has only a definite volume but no definite shape. A gas has neither a fixed volume nor shape. What is responsible for this difference in the properties of the three states of matter ?

To seek an answer to this question, we have to find out what matter is made up of. Let us first examine the behaviour of gases.

Diffusion

When a bottle of perfume is opened the pleasant smell is felt in every part of the room. How does this happen ? The perfume may have become vapour and somehow managed to spread itself throughout the available space in the room. Similarly when a small amount of acetic acid (vinegar) is taken in a dish and left in a corner of the room, acetic acid vapours also spread throughout the room in a few minutes. When a gas jar is inverted over another containing brown nitrogen dioxide, the gas moves slowly into the upper jar and fills it completely.

These observations suggest that all gases are capable of moving from one place to another and tend to occupy the entire space available to them. *This phenomenon is known as diffusion.*

The movement of a gas in space can be explained by assuming that all gases are made up of small particles which move rapidly in all directions and occupy all the available space. These particles are too small to be seen.

How does one gas move into another ? We can reasonably assume that there are spaces between the particles of gases. The mixing of gases must, therefore, be due to empty spaces between the particles and the tendency of the particles to be in motion.

Do particles of all gases move with the same speed ? When two cotton plugs, one soaked in concentrated hydrochloric acid and the other in liquor ammonia, are introduced at the two ends of a glass tube mounted horizontally, we notice the formation of a white deposit in the form of a ring (Fig. 2.1).

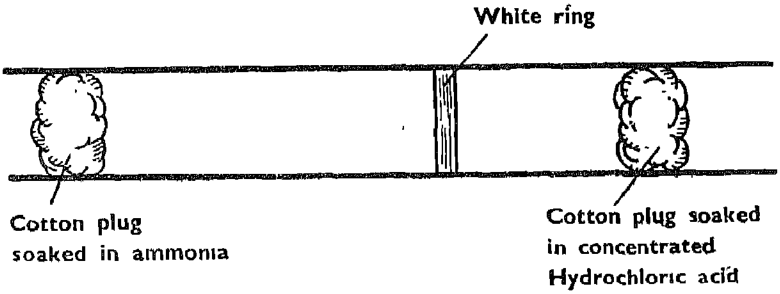


Figure 2.1 *Relative diffusion of ammonia and hydrogen chloride*

The ring is found to be closer to the cotton plug soaked in hydrochloric acid than to that soaked in ammonia. This shows that ammonia moves faster than hydrogen chloride. Thus the particles of ammonia and hydrogen chloride move with different speeds.

Why is this so ? The speed of the particles depends on their weight, heavier particles moving slower and the lighter ones faster. As ammonia gas moves faster than hydrogen chloride, the particles of ammonia must be lighter than those of hydrogen chloride.

Gases exert pressure

When air is blown into a rubber balloon, it bulges out on all sides. It becomes bigger as more and more air is blown into it.

Why does the balloon bulge out ? The particles present in the air blown into the balloon move rapidly in all directions. When they do so, they collide with one another and also strike and push the walls of the balloon and rebound. Thus pressure is developed when the particles strike the walls. Hence the balloon bulges out. Since the particles are uniformly distributed, they hit the walls of the balloon equally on all sides. Pressure exerted on the walls of the balloon is therefore equal in all directions.

Effect of pressure on the volume of a gas at a given temperature

What happens to the volume of a gas if its pressure is changed ? Let us look at a small volume of air confined over water in a manometer closed at one end and bent at right angles at the other. The bent side is connected to a syringe through a rubber tubing (Fig. 2.2).

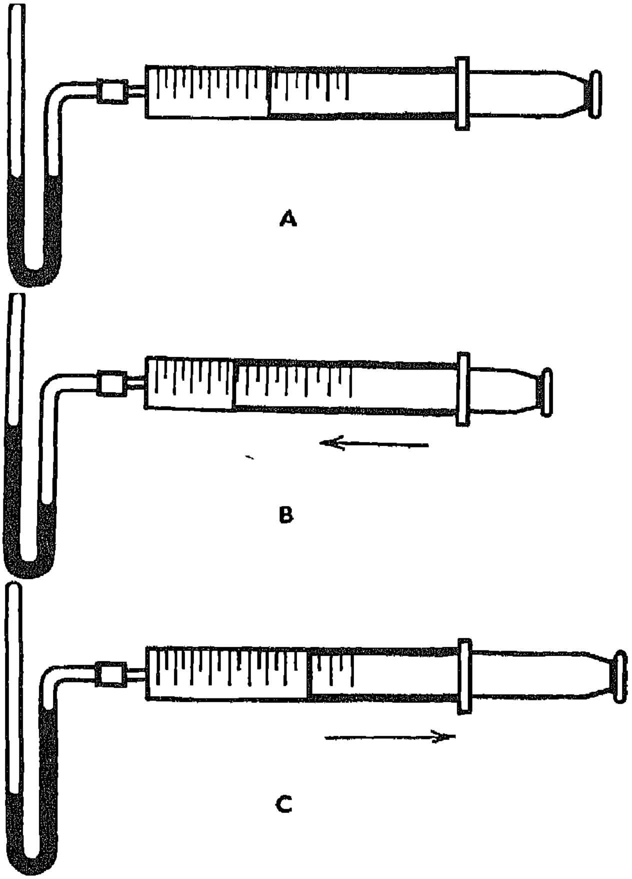


Figure 2.2 *Effect of pressure on the volume of a gas*

Air occupies a certain volume for a particular position of the piston in the syringe (A). If the pressure on the air is increased by pushing the piston forward, the water level in the closed limb rises. The volume occupied by air is now reduced (B). So the volume of air is decreased when pressure is increased. If the pressure is decreased by drawing the piston backwards, the level of water in the closed limb falls (C). The volume occupied by air increases. So, on decreasing the pressure, the volume of air is increased.

These observations could be explained by assuming that gases contain particles. Decrease of volume with increase in pressure suggests that the particles are far apart and are brought closer by increasing pressure. When the pressure is decreased, the particles move farther apart. The distances between them increase. The volume of the gas, therefore, increases.

Effect of temperature on the volume of a gas at constant pressure

When a boiling tube fitted with a capillary containing a mercury seal is heated, the mercury seal rises in the capillary (Fig. 2.3).

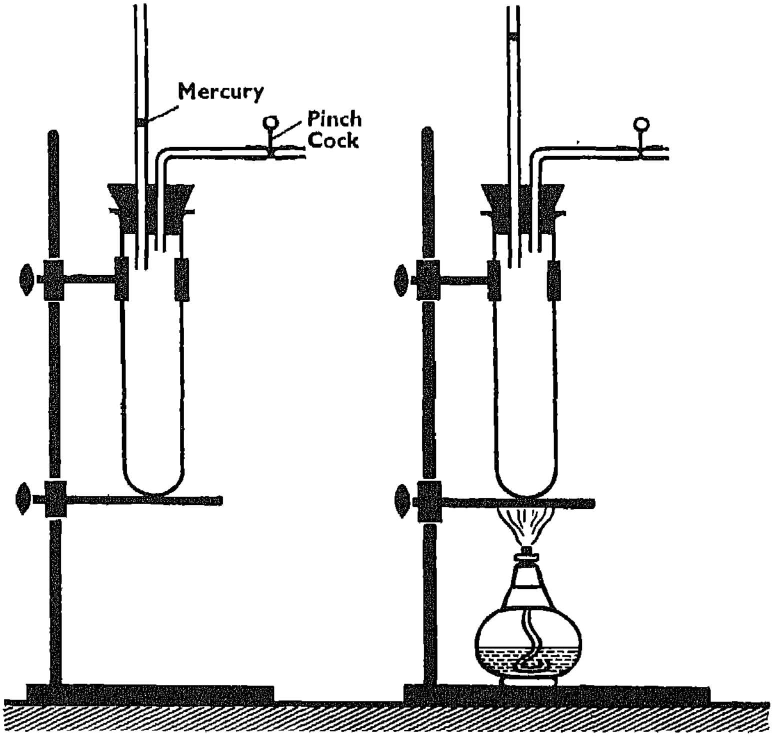


Fig. 2·3 *Expansion of air on heating.*

On cooling, the mercury comes down. This shows that the air in the boiling tube expands on heating and contracts on cooling.

When the temperature is raised, the particles of air move apart. The distances between them increase. Consequently the volume of the air increases. If however, the temperature is reduced, the particles come closer. So the volume of air decreases.

It is now reasonable to suppose that a gas contains particles. These particles are called **molecules**. From what we have studied so far we learn that

(i) molecules are exceedingly small in size

(ii) there are spaces between molecules

(iii) molecules are in motion

(iv) molecules move with different speeds

(v) molecules of different substances are different.

2.2 Nature of solids and liquids

We learnt in unit 1 (page 11) that steam (gas) on cooling, becomes water (liquid). Water on further cooling, becomes ice (solid). As the liquid and solid are obtained from the gas, the molecules present in the gas must also be present in the liquid and in the solid. Water and ice, therefore, contain the same molecules which are present in steam. If ice and water are made up of the same molecules why do they have different properties? There must be some difference in the arrangement of molecules in the solid and liquid states.

A small amount of water, if allowed to boil in a closed kettle vapourises and the steam formed occupies the whole kettle. If this steam is condensed, it gives back water. It is obvious that the small amount of water in the form of liquid has turned into a larger volume of steam. Therefore, a given mass of a substance occupies a relatively smaller volume in the liquid state and a much larger volume in the gaseous state. So the molecules in water are closer to one another than they are in steam.

When a crystalline solid (for example, sulphur) is melted it becómes mobile. That is, the rigidity of the solid state is lost. In most cases, it is found that liquid occupies a slightly larger volume than the solid. It therefore appears that the molecules are a little farther apart in the liquid and arranged in a less orderly manner than in the solid.

The manner of arrangement of molecules in the three states of matter is schematically given in Fig. 2.4.

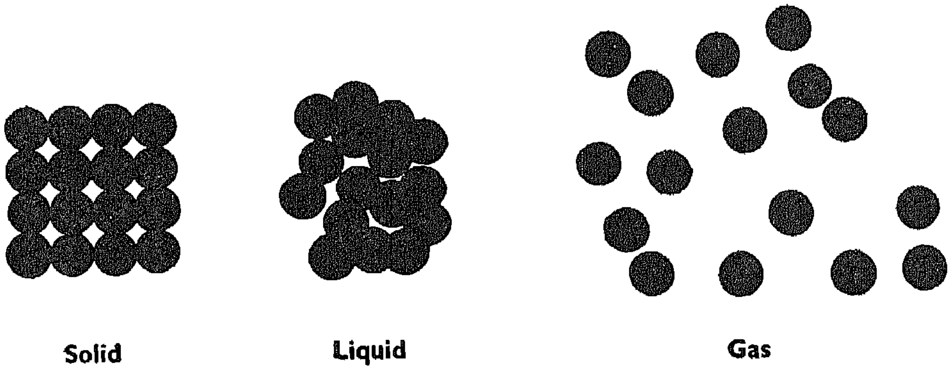


Figure 2.4 *Molecules in the three states of matter*

Molecules attract one another Forces of attraction depend upon the distances between the molecules. The smaller the distance, the stronger will be the forces of attraction. As the molecules in a solid are very close to one another it is reasonable to assume that the forces of attraction between them are great. As a result, the molecules have little freedom of movement. Thereby they remain fixed in space thus giving a solid rigidity and definite shape.

In a liquid, where the molecules are slightly farther apart, the binding forces are not so great. There are small empty spaces between the neighbouring molecules. The weaker attractive forces enable the molecules to slide into the empty spaces thereby making the liquid mobile. For this reason, liquids do not possess a definite shape, but simply assume the shape of the container.

In a gas on the other hand, the molecules are far away from one another. The attractive forces are therefore extremely weak. Hence the molecules have by far the greatest freedom of movement. They occupy the entire space available.

Thus we find that *the attractive forces, the distances between molecules and their arrangement satisfactorily account for the differences in the behaviour of solids, liquids and gases,*

2.3 Change of state

Is it possible to explain the change of one state of matter into another in terms of the distances and attractive forces between the molecules ? Change of solid to liquid and liquid to gas involves overcoming the attractive forces and separation of molecules by larger distances. Energy is required for overcoming the attractive forces and is supplied in the form of heat.

When a solid is heated, the molecules in it gain more energy. They overcome the forces of attraction and move away from one another. The molecules are now in a more disorderly state which enables them to slip past one another. A liquid therefore results.

When the liquid is heated the molecules pick up some more energy and move even farther apart. Now they have a greater freedom of motion and can move in all directions. Thus a gaseous state results.

On cooling the gas, the molecules of the gas lose energy and are brought nearer to each other by attractive forces, resulting in a liquid. When the liquid is further cooled, the molecules lose more energy, come even closer together resulting in a solid.

2.4 Atoms and molecules

We learnt so far that solids, liquids and gases contain molecules. Are the molecules the smallest particles of matter or do they contain still smaller particles ? When water is electrolysed, hydrogen and oxygen are obtained in the ratio 2 : 1 by volume. It is also found that when a mixture of two volumes of hydrogen and one volume of oxygen is taken and sparked, the two gases combine completely forming water. This shows that hydrogen and oxygen combine in the ratio 2 : 1 by volume.

$$\underset{(2\text{ vols.})}{\text{Hydrogen}} + \underset{(1\text{ vol.})}{\text{Oxygen}} \longrightarrow \text{Water}$$

Amadeo Avogadro put forward a hypothesis relating the volume occupied by a gas to the number of molecules present in it. According to him *equal volumes of gases contain equal number of molecules at the same temperature and pressure*, whatever be the nature of the gas. If one volume of oxygen contains, say, a million molecules, then one volume of any other gas (for example, hydrogen, nitrogen or carbon dioxide) should also contain the same number (one million) of molecules. Two volumes of a gas should contain 2 million molecules and 3 volumes should contain 3 million molecules.

In the formation of water, 2 volumes of hydrogen combine with 1 volume of oxygen. If there are 1 million molecules in oxygen then, according to Avogadro, there must be 2 million molecules in hydrogen.

Hydrogen + Oxygen —→Water

2 million moleculesof hydrogen + 1 million molecules of oxygen —→Water

or

2 molecules of hydrogen + 1 molecule of oxygen—→Water

or

1 molecule of hydrogen + $\frac{1}{2}$ molecule ofoxygen —→Water

So each molecule of hydrogen should combine with half a molecule of oxygen. This means that a molecule of oxygen can be split into two fragments. The smaller fragments obtained from the molecules are called **atoms**.

Atoms are known to be very small. They cannot be seen even under a very powerful microscope. They have a definite weight, although very small. Atoms of one element are different from atoms of another element. Unlike molecules, atoms generally do not exist free. Atoms are very reactive. Two or more atoms combine to form molecules. The combining atoms may be of the same type or of different types. These ideas were first developed by **John Dalton**. Hydrogen, oxygen, nitrogen and bromine which contain atoms of the same kind are called elements. A molecule of water contains 2 atoms of hydrogen and 1 atom of oxygen. A molecule of carbon dioxide contains 1 atom of carbon and 2 atoms of oxygen. A molecule of ammonia contains 3 atoms of hydrogen and 1 atom of nitrogen (Fig. 2.5).

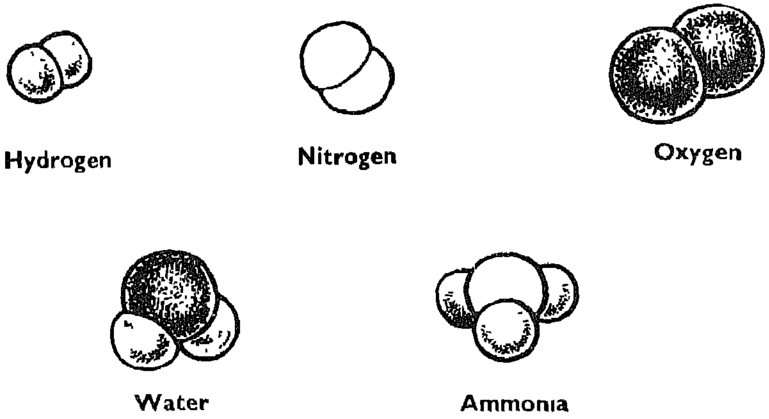


Figure. 2.5 *Models of some simple molecules.*

As water, ammonia and carbon dioxide contain more than one kind of atoms they are called compounds. It is possible to explain how two volumes of hydrogen combine with one volume of oxygen, producing two volumes of steam (Fig. 2.6).

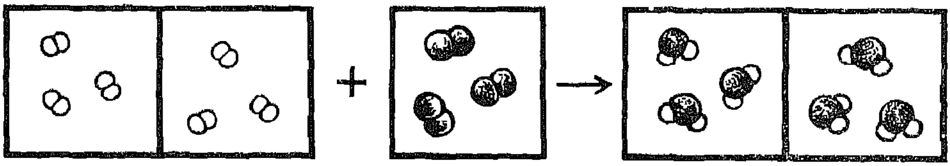


Figure 2.6 *Combination of hydrogen and oxygen to form steam*

2.5 Modern picture of the atom

For quite a long time atom was believed to be the smallest particle of matter. Is it really the smallest particle of matter, or are there any still smaller particles in it?

It has been shown that a plastic comb, when rubbed with a silk cloth, picks up bits of paper and attracts a suspended pith ball and a thin stream of water from a burette. Before rubbing, it does not show these properties. What has happened to the comb on rubbing?

When the comb is rubbed with silk, it acquires a charge. Some charge must have been transferred from silk to the comb leaving silk oppositely charged. Opposite charges attract each other. Matter, therefore, is electrical in nature. Scientists found that an atom can be split into small charged particles. The negatively charged particles are called **electrons**. The positively charged particles are called **protons**. A proton and an electron carry equal amounts of charge but of opposite type. There are also particles without charge. These are called **neutrons**. An atom is made up of these three types of particles. You will learn more about this in higher classes. As an atom is electrically neutral, the number of negatively charged particles (electrons) must be the same as the number of positively charged particles (protons). The protons and neutrons occupy a very small space at the centre of the atom called **nucleus**. The electrons are distributed in space around the'nucleus. So an atom is no longer considered the smallest particle of matter. It consists of still smaller particles such as electrons, protons and neutrons.

2.6 How heavy are atoms and molecules ?

We have learnt that atoms are exceedingly tiny particles. The weights of such particles must be extremely small. It is very difficult to weigh them. Instead of actual weights, therefore, it is convenient to use comparative weights. The weight of carbon atom is taken as 12 and all other atoms of elements are compared with it. The atomic weight of hydrogen on this scale is 1.008 which for all purposes is taken as 1.0.

The comparative weights of atoms are known as atomic weights. Atomic weight of an element is therefore not the actual weight of an atom of the element, but shows how many times its atom is heavier than that of hydrogen. For example, the atomic weight of oxygen is 16. This means that an atom of oxygen is sixteen times heavier than an atom of hydrogen. If the atomic weight is expressed in grams, it is called *gram-atomic weight*. Gram-atomic weight of oxygen is 16.0 g. It is experimentally found that one gram-atomic weight of any element contains 6.0 x 10^{23} atoms. This number is known as **Avogadro's number**.

Table 1 Atomic weights of some common elements

Element	Atomic Weight
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Aluminium	27
Sulphur	32
Iron	56
Zinc	65
Lead	207

2.7 Molecular weight*

Since a molecule is composed of atoms, the weight of a molecule must be the sum of the weights of atoms present in it. For example, a molecule of water contains two atoms of hydrogen and one atom of oxygen. The weight of a molecule of water will be the sum of the weights of 2 atoms of hydrogen and one atom of oxygen. If the actual weights of the atoms are used, the weight of the molecule will be very small. So atomic weights are used instead of actual weights of atoms The sum of the atomic weights will give the molecular weight. The molecular weight of water will then be (2 x 1) + (1 x 16) = 18, where 1 and 16 are the atomic weights of hydrogen and oxygen respectively.

Molecular weight is also a comparative weight. It is the ratio of the weight of a molecule of a substance and the weight of an atom of hydrogen.

$$\text{Molecular weight} = \frac{\text{Weight of 1 molecule of a substance}}{\text{Weight of 1 atom of hydrogen}}$$

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*In the case of ionic solids formula weight is used instead of molecular weight

Molecular weight is, therefore, only a number showing how many times a molecule is heavier than an atom of hydrogen. The molecular weight of water is 18. Hence a molecule of water is 18 times heavier than an atom of hydrogen. Similarly, a molecule of carbon dioxide (mol. wt. 44) is 44 times heavier than one atom of hydrogen.

If the molecular weight is expressed in grams, it is called *gram-molecular weight*. The gram-molecular weight of water is 18 g. It contains 6.0×10^{23} (Avogadro's number) molecules of water. Avogadro's number stands for the number of molecules present in one gram-molecular weight of any element or compound.

The term 'mole' is now used to represent a gram-atomic weight of an element or a gram-molecular weight of a substance.

Questions

1. How do gases diffuse ?
2. Do all gases diffuse at the same speed ? Give reasons.
3. Why does a balloon bulge when air is blown into it ?
4. What does a gas consist of ? Give your arguments in support of your answer.
5. What happens to the volume of a gas when
   a. pressure is increased
   b. temperature is increased ?
6. Are there spaces between molecules in a gas ? If so, what is the evidence ?
7. What do solids and liquids consist of ?
8. What are the essential differences between solids, liquids and gases ?
9. Explain why solids have a definite shape and liquids do not.
10. What is a molecule made up of ?
11. How many atoms are there in a molecule of
    a. hydrogen
    b. water and
    c. carbon dioxide ?
12. Are the molecules of ammonia and carbon dioxide similar ?

13. Is atom the smallest particle of matter ?

14. How many atoms are there in a gram-atomic weight of oxygen ?

15. What is atomic weight ?

16. What is Avodgadro's number ?

17. Why is atom electrically neutral, even though it contains charged particles ?

18. How is gram-atomic weight related to atomic weight ?

19. Fill up the blanks with suitable words.

    a. Atoms of the same element are.. ....
    atoms of different elements are...... ....

    b. Molecules of elements contain atoms of the. . . . . kind while molecules of compounds contain atoms of . . .. kinds.

    c. The number of atoms present in a gram-atom of the element is called... . . . .

    d. Gases contain . ... which move in . . . .directions with . .. . speeds.

    e. Molecules in a gas strike the walls and develop. ....

    f. At constant volume, the pressure of a given mass of a gas . ... if its temperature is raised.

    g. At constant temperature, the volume of a given mass of gas .. . . as its pressure is increased.

    h. Two gases of the same volume at same temperature and pressure contain. .. . . .number of molecules.

    i. If one volume of oxygen contains 10,000 molecules one volume of hydrogen should contain .. . . .molecules (at a given temperature and pressure).

20. Fill up the blanks choosing the correct word.

    a. Particles in a lighter gas move... .. ..(slower, faster) than those in a heavier one.

    b. Increase of pressure. .... (increases, decreases) the volume of a given mass of gas at a constant temperature.

    c. Volume of a given mass of gas .... (increases, decreases) when its temperature is decreased at a constant pressure

    d. The molecules of a gas move ....(slower, faster) at a higher temperature than at a lower temperature.

e. A gas occupies. . . . .(all, only a part of), the space available, to it.

f. Different gases move with ... . (same, different) speeds

g. Pressure exerted by a gas is . . . .(equal, unequal) in all directions

h. Gas molecules are at relatively.. .. . .. (large, small) distances from each other.

21. Match the following

	A	B
i)	A solid	has no definite shape.
ii)	A gas	has definite shape
iii)	A liquid	occupies all the available space
iv)	A solution	contains only one kind of matter.
v)	An element	is a homogeneous mixture.

22. State whether the following statements are TRUE or FALSE.

a. Atoms are not subdivided during a chemical reaction.

b. Atom is the smallest particle of matter

c. Number of protons in an atom is equal to the number of electrons.

d. Atom is electrically neutral.

e. Two electrons repel each other and two protons also repel each other

f. Elements contain atoms of the same kind while compounds contain atoms of different kinds.

g. The volume of a gas decreases when pressure is increased.

h. Molecules of all gases move with the same speed

UNIT 3

Symbols, Formulae and Equations

To represent the changes that matter undergoes, chemists have evolved a language of their own. Elements and compounds are represented by symbols and formulae and chemical changes by equations. In this unit we shall learn how to write formulae and equations and make use of them in calculations

3.1 Symbols of elements

In unit 1 (page 11) we learnt about the changes that matter undergoes. For example, a piece of magnesium, when heated in air, burns with a dazzling white light, and changes to a white ash (magnesium oxide). In this reaction magnesium combines with oxygen of the air to form magnesium oxide. To express such changes, we may use chemical names of the substances involved in the reaction. We can write the above change as

$$\text{Magnesium}+\text{Oxygen}\longrightarrow \text{Magnesium oxide.}$$

The names of substances are usually long. So, for the sake of convenience, they are written in a short form using some **symbols.** The use of symbols is not peculiar to chemistry alone, but is a common practice. We use, for example, the symbols + and - to express addition and subtraction.

Substances involved in a chemical reaction may be elements or compounds. For representing elements, **Berzelius** suggested a simple system. The initial letter (in capital) of the English name of the element is used to represent the element.

Element	**Symbol**
Hydrogen	H
Oxygen	O
Carbon	C
Nitrogen	N
Iodine	I
Sulphur	S

What should we do if several elements have names starting with the same letter? For example, carbon, chlorine, and calcium have C as the first letter; magnesium and manganese have M. In such cases, we use either the first two letters or the first and some other letter of the name as symbol of the element. The first letter is always a capital.

Element	**Symbol**
Carbon	C
Calcium	Ca
Chlorine	Cl
Magnesium	Mg
Manganese	Mn

In some cases, symbols are derived from the Latin names of the elements.

Element	**Latin Name**	**Symbol**
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Sodium	Natrium	Na
Potassium	Kalium	K
Silver	Argentum	Ag
Gold	Aurum	Au

The symbols of a few common elements, together with their atomic weights are given below:

Element	**Symbol**	**Atomic weight**
Aluminum	Al	27
Calcium	Ca	40

Element	Symbol	Atomic weight
Carbon	C	12
Chlorine	Cl	35.5
Copper (Cuprum)	Cu	63.5
Gold (Aurum)	Au	197
Hydrogen	H	1
Iodine	I	127
Iron (Ferrum)	Fe	56
Lead (Plumbum)	Pb	207
Magnesium	Mg	24
Manganese	Mn	55
Mercury(Hydrargyrum)	Hg	201
Nickel	Ni	59
Nitrogen	N	14
Oxygen	O	16
Phosphorus	P	31
Potassium (Kalium)	K	39
Silicon	Si	28
Silver (Argentum)	Ag	108
Sodium (Natrium)	Na	23
Sulphur	S	32
Tin (Stannum)	Sn	119
Zinc	Zn	65

3.2 Meaning of a symbol

The symbol of an element is not only an abbreviation of its name, but also represents one atom as well as one gram–atom of the element. Thus, the symbol H represents one atom and one gram-atom of hydrogen. It may be recalled here that a gram–atom of an element contains 6.0×10^{23} atoms.

3.3 Formulae of molecules

Just as symbols are used to represent atoms, **formulae** are used to represent molecules. A molecule may contain two or more atoms of the same element or of different elements. A molecule of hydrogen for example, contains two atoms. Its molecular formula is therefore written as H_2. The subscript 2 indicates the number of atoms present in a molecule of hydrogen. Similarly molecules of oxygen, nitrogen and chlorine are represented by O_2 , N_2 and Cl_2 respectively as they contain two atoms each. A molecule of ozone consists of three atoms of oxygen and so its molecular formula is O_3.

A molecule of hydrogen chloride contains one atom of hydrogen and one atom of chlorine. It is represented by writing the symbols of hydrogen and chlorine side by side as HCl. When a molecule contains more than one atom of any kind, the number is indicated by a subscript at the lower right hand corner of the symbol For example, a molecule of water contains two atoms of hydrogen and one atom of oxygen. Its formula is therefore written as H_2O, the subscript 2 indicating the number of hydrogen atoms present in a molecule of water. Similarly a molecule of carbon dioxide is represented by CO_2, as its molecule contains one atom of carbon and two atoms of oxygen.

To represent more than one molecule of a substance, we write the appropriate number before its formula. For example, two molecules of water are represented by $2H_2O$ and three molecules of carbon dioxide by $3CO_2$.

Significance of a formula

What does a formula tell us? The formula of carbon dioxide, CO_2, gives the following information·

1. Carbon dioxide contains the elements carbon (C) and oxygen (O).
2. Each molecule of carbon dioxide contains one atom of carbon and two atoms of oxygen.
3. In chemical calculations it stands for 1 gram-molecule of carbon dioxide. A gram-molecule contains 6.0×10^{23} molecules.
4. It helps in the calculation of formula weight. The formula weight of CO_2 is given by the sum of the atomic weight of carbon and twice the atomic weight of oxygen.

$$C + 2O = CO_2$$

$$12 + 2 \times 16 = 44.$$

Percentage composition from a formula

The weight of each element present in 100 g of a compound is the percentage weight of the element.

Let us calculate the percentage weight of carbon and oxygen in carbon dioxide. The formula of carbon dioxide is CO_2. Its formula weight is

$C+2O=12+2\times16=44$

44g of CO_2 contain 12 g of carbon

100 g of CO_2 contain $\frac{100\times12}{44}=27.3$ g of carbon

Hence percentage by weight of carbon in $CO_2=27.3$

Similarly percentage of oxygen is $\frac{100\times32}{44}=72.7$

Formula from the percentage composition

We know how to write the formula of a compound if we know the number of atoms of different elements present in a molecule of that compound. This information is obtained from the percentage composition and molecular weight of the substance.

Let us consider a compound having 27.3% by weight of carbon and 72.7% of oxygen. The ratio of atoms in this compound is calculated as follows :

Element	Symbol	%of the element in the compound	Atomic weight of the element	$\frac{\text{Percentage}}{\text{Atomic weight}}$
Carbon	C	27.3	12	$\frac{27.3}{12}=22.7$
Oxygen	O	72.7	16	$\frac{72.7}{16}=4.54$

The ratio of carbon to oxygen atoms is, therefore,

$$C:O=2.27:4.54$$
$$=1:2$$

From this ratio, the formula of the compound may be written as CO_2.

The molecular formula of a compound

The calculation of the ratio of atoms in the example just given does not necessarily give the formula of a molecule of carbon dioxide, since the ratio of one atom of carbon to two atoms of oxygen is satisfied by the formulae CO_2, C_2O_4, C_3O_6 or some other higher multiple of CO_2. The choice between these is made on the basis of the molecular weight. The molecular weights corresponding to the formulae written above would be

for CO_2	for C_2O_4	for C_3O_6
$C=1\times 12 = 12$	$2\ C=2\times 12=24$	$3\ C=3\times 12=36$
$2\ O=2\times 16 = 32$	$4\ O=4\times 16=64$	$6\ O=6\times 16=96$
44	88	132

Experimentally the molecular weight of carbon dioxide is found to be 44. This weight corresponds to the formula CO_2. Hence CO_2 represents the molecular formula of carbon dioxide. A formula obtained in this way is called the molecular formula, since it truly represents the formula of a molecule of the compound.

The formulae and formula weights of some common substances are given below :

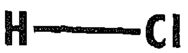
Substance	**Formula**	**Formula Weight**
Calcium carbonate	$CaCO_3$	100
Carbon dioxide	CO_2	44
Cupric oxide	CuO	79.5
Hydrochloric acid	HCl	36.5
Magnesium oxide	MgO	40.0
Nitric acid	HNO_3	63
Potassium nitrate	KNO_3	101
Sodium chloride	$NaCl$	58.5
Sodium hydroxide	$NaOH$	40
Sulphuric acid	H_2SO_4	98
Water	H_2O	18

3.4 Valency and chemical formulae

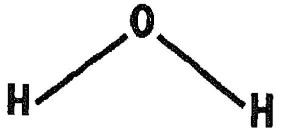
The formula of a compound can be written by making use of the combining capacity of the elements present in it. The elements can be present either as atoms or ions.

Valency of atoms

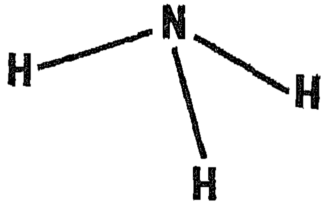
It is known that one atom of chlorine combines with one atom of hydrogen to form one molecule of hydrogen chloride. One atom of oxygen on the other hand, combines with two atoms of hydrogen to form a molecule of water. One atom of nitrogen combines with three atoms of hydrogen to form a molecule of ammonia. Thus one atom each of chlorine, oxygen and nitrogen combines with one, two and three atoms of hydrogen respectively. The number of hydrogen atoms with which one atom of an element combines is termed its combining capacity. Thus the combining capacities of chlorine, oxygen and nitrogen are one, two and three respectively. *The combining capacity of an element is called* **valency**.



Hydrogen Chloride



Water



Ammonia

In the case of those elements which do not combine with hydrogen, their valencies are determined with reference to chlorine, oxygen or nitrogen, which have valencies 1, 2 and 3 respectively. The valency of magnesium is determined with respect to oxygen. One atom of magnesium combines with one atom of oxygen which, in turn, combines with two atoms of hydrogen. Hence the valency of magnesium is two.

Charge of ions

It was mentioned in unit 2 (page 27) that an atom consists of positively charged nucleus surrounded by negatively charged electrons and that the number of electrons in an atom is equal to the number of protons.

Atoms lose or gain electrons. When an atom loses an electron the balance between positive and negative charges is upset, there being one positive charge in excess. The atom is no longer neutral, but has a unit positive charge. Similarly when an atom gains an electron, it acquires a unit negative charge.

Atoms or groups of atoms carrying positive or negative charge (one or more units depending on the number of electrons lost or gained) are called **ions**. The amount and type of charge carried by an ion are indicated by the appropriate number of positive or negative signs written on the right upper side of the symbol. The combining capacity (valency) of an ion is equal to its positive or negative charge. For example, when sodium reacts with chlorine, each sodium atom loses an electron and forms a positively charged sodium ion, Na^+. This electron is gained by a chlorine atom, which forms a negatively charged chloride ion, Cl^-. Thus sodium chloride (Na^+Cl^-) contains positive and negative ions. Such compounds are called ionic compounds.

The symbols and charges of some common ions are given below.

Ion	**Symbol**	**Charge**
Aluminium	Al^{+++}	+3
Ammonium	NH_4^+	+1
Calcium	Ca^{++}	+2
Carbonate	CO_3^{--}	–2
Chloride	Cl^-	–1
Copper (I) (Cuprous)	Cu^+	+1
Copper (II) (Cupric)	Cu^{++}	+2
Hydrogen	H^+	+1

Ion	**Symbol**	**Charge**
Hydroxide	OH^-	-1
Iodide	I^-	-1
Iron (II) (Ferrous)	Fe^{++}	+2
Iron (III) (Ferric)	Fe^{+++}	+3
Lead	Pb^{++}	+2
Magnesium	Mg^{++}	+2
Mercury (II)	Hg^{++}	+2
Nickel	Ni^{++}	+2
Nitrate	NO_3^-	-1
Oxide	O^{--}	-2
Potassium	K^+	+1
Silver	Ag^+	+1
Sodium	Na^+	+1
Sulphate	SO_4^{--}	-2
Sulphide	S^{--}	-2

Formula of an ionic compound

Although an ionic compound contains positive and negative ions, it contains no excess positive or negative charge. The number of positive and negative ions present in an ionic compound is such that the total positive charge is equal to the total negative charge.

Let us take the case of sodium sulphate. It consists of sodium ions, (Na^+) and sulphate ions, (SO_4^{--}). To keep the positive and negative charges balanced, two sodium ions and one sulphate ion must be present.

While writing the formula of sodium sulphate, the positive ion (Na^+) is written first, followed by the negative ion (SO_4^{--}). The number of ions of each kind is indicated by a subscript. The formula of sodium sulphate is therefore written as Na_2SO_4, omitting the charges of the ions.

$$\begin{matrix} Na^+ \\ \\ Na^+ \end{matrix} + SO_4^{--} \longrightarrow Na_2SO_4$$

What is the formula of ferric oxide?

Ferric oxide contains ferric ions, Fe^{+++} and oxide ions, O^{--}. To balance the charges, two ferric ions must combine with three oxide ions to form ferric oxide,

$$\begin{array}{lcll} Fe^{+++} & & O^{--} & \\ & + & & O^{--} \longrightarrow Fe_2O_3 \\ Fe^{+++} & & O^{--} & \end{array}$$

In order to write the formula of an ionic compound, therefore, it is necessary to know:

1. the ions from which it is formed
2. the symbols or the formulae of the ions and
3. the valencies or the number of charges of the ions.

3.5 Chemical equations.

The formation of magnesium oxide, when magnesium burns in oxygen is represented as

$$\text{Magnesium} + \text{Oxygen} \longrightarrow \text{Magnesium oxide} \qquad (1)$$

The substances taking part in the reaction (reactants), namely magnesium and oxygen are written on the left. The substance(s) formed during the reaction (products), namely magnesium oxide is written on the right. The reactants and products are connected by an arrow. The arrow head points towards the products.

Using symbols and formulae, the reaction between magnesium and oxygen could be written as

$$\begin{array}{ccccc} Mg & + & O_2 & \longrightarrow & MgO \qquad (2) \\ \text{Magnesium} & & \text{Oxygen} & & \text{Magnesium oxide} \end{array}$$

In a chemical change the mass of the reactants is equal to the mass of the products (page 16). So the number of atoms of each element present on either side of the equation must be balanced.

In equation (2) there are two atoms of oxygen on the left, while there is only one on the right. This equation is not correct. It has to be modified to equalise the number of atoms of oxygen and magnesium on either side The modified equation is

$$2\ Mg + O_2 \longrightarrow 2\ MgO \qquad (3)$$

Equation 3 is called a *balanced equation* because the number of atoms of each element on the left balance the number of atoms of each element on the right.

Let us now write the equation for the formation of water from hydrogen and oxygen in words.

$$\text{Hydrogen} + \text{Oxygen} \longrightarrow \text{Water} \qquad (4)$$

Hydrogen and oxygen are diatomic molecules. They are represented by H_2 and O_2. Thus equation (4) becomes

$$H_2 + O_2 \longrightarrow H_2O \qquad (5)$$

This equation is not balanced with respect to oxygen. To balance the oxygen atoms, we multiply H_2O by 2

$$H_2 + O_2 \longrightarrow 2H_2O \qquad (6)$$

In so doing, we have increased the number of hydrogen atoms from 2 to 4. So the number of hydrogen atoms on the left should be increased to 4. This is done by multiplying H_2 by 2. Thus we get

$$2H_2 + O_2 \longrightarrow 2H_2O \qquad (7)$$

This is a balanced equation because the number of atoms of hydrogen and oxygen on both sides is the same.

The balanced equation for the burning of carbon in oxygen forming carbon dioxide gas can be represented by

$$C + O_2 \longrightarrow CO_2 \uparrow \qquad (8)$$

Here the product formed is a gas. In such cases the evolution of a gas is indicated by an arrow pointing upwards.

When a solution of sodium chloride is added to a solution of silver nitrate, a white solid (precipitate) of silver chloride is thrown down and the balanced equation is represented by

$$Na\,Cl + Ag\,NO_3 \longrightarrow Ag\,Cl \downarrow + Na\,NO_3 \qquad (9)$$

In such reactions, the formation of a precipitate as one of the products is indicated by an arrow pointing downwards.

A chemical equation is a short and simple representation of a chemical reaction. An equation for a chemical reaction can be written if the formulae of the reactants and the products are known.

Questions

1. Write the chemical symbols for oxygen, hydrogen, nitrogen, chlorine, sulphur, carbon, copper, iron, silver, gold, calcium and magnesium.
2. Name the elements which have the following symbols · K, Na, Au, Ag, Ba, Ca, Zn, S, P and Hg.
3. Several elements have the same initial letter. How do we indicate these elements by symbols?
4. Name two symbols derived from the Latin names of the elements concerned.
5. Explain the meaning of the formulae HCl and NaCl.
6. How many atoms of the various elements are indicated by each of the following formulae? $2H_2O$, 5HCl, $7HNO_3$.
7. The formula for sulphuric acid is H_2SO_4.
   a. What elements are present in this compound?
   b. How many atoms of each element are present in one molecule of this compound?
   c. What is the formula weight of the compound?
8. Calculate the formula weights of the following compounds.
   a. Potassium chloride, KCl
   b. Zinc oxide, ZnO
   c. Calcium chloride, $CaCl_2$
9. Find the percentage composition of the following compounds.
   a. NH_3 b. CaC_2 c. Fe_2O_3
10. Write the symbols of the following with their charges.
    a. Magnesium ion b. Hydroxide ion
    c. Chloride ion d. Sulphate ion

11. Write the formulae of the compounds formed in the following cases.

Ions	Cl^- Chloride	NO_3^- Nitrate	OH^- Hydroxide	S^{--} Sulphide	O^{--} Oxide	SO_4^{--} Sulphate	CO_3^{--} Carbonate
Ag^+							
Mg^{++}							
Al^{+++}							

12. Which ions are present in the following

a. NaOH, b. CuO, c. $ZnSO_4$ d. $CaCl_2$?

13. Correct the following formulae wherever necessary (Use table 3.3).

$MgCl$, Ag_2O, KSO_4, $NaCl$, $MgSO_4$, AlO_2, $CaCl_3$ and $MgOH$.

14. What facts must be known before a chemical equation can be written?

15. Fill up the blanks with suitable words.

a. Gram-atom of an element contains . . . . atoms

b. The composition of the molecule of a substance is represented by a chemical. . . . .

c A gaseous element involved in an equation is usually written as a molecule consisting of two . . . .

d. An equation is balanced to make it conform to the law of. . . . .

e The formula weight of a compound is the sum of the . . . . of the elements present in it.

f. Just as symbols are used to represent atoms, formulae are used to represent . . . .

16. Against each symbol in group B, write the number corresponding to the name of the element in group A.

GROUP A

1. Calcium
2. Chlorine
3. Copper
4. Sodium
5. Magnesium
6. Lead
7. Manganese
8. Tin
9. Potassium

GROUP B

Cu . . . .
K. . . . .
Ca . . . .
Mg . . . .
Pb . . . .
Na . . . .
Mn. . . . .

17 Balance the following equations:

a. $Al + Cl \longrightarrow AlCl_3$

b. $KClO_3 \longrightarrow KCl + O_2$

c. $N_2 + H_2 \longrightarrow NH_3$

d. $Mg + N \longrightarrow Mg_3N_2$

18 (i) Deduce the formula of a compound having the following composition by weight:

a Ca=29.4% S=23.5% O=47.1% and

b. Cu=39.5% S=20.3% O=40.2%

(ii) Calculate the masses (weights) of the following:

a. 1 mole of CO_2

b 0.5 mole of H_2O

c. 0.1 mole of NH_3

d. 2.5 moles of $CaCO_3$

(iii) Calculate the number of moles in the following:

a. 1000 g of water (H_2O)

b. 34.0 g of silver nitrate ($AgNO_3$)

c. 31.6 g of potassium permanganate ($KMnO_4$)

(iv) Calculate the number of molecules present in one gram of calcium carbonate ($CaCO_3$).

UNIT 4

Solutions

Solutions play a vital role in life processes. Many chemical reactions take place in solution. In this unit we shall learn what a solution is, the factors affecting the rate of solution and the heat changes accompanying dissolution.

4.1 What is a solution ?

When a spoonful of sugar is added to a glass of water and stirred, the sugar disappears The resulting liquid tastes like sugar. Sugar is said to have dissolved in water and the liquid is called a solution of sugar in water.

Many other substances like common salt, copper sulphate and potassium permanganate dissolve in water to form solutions. Most of the common solutions are formed by dissolving a solid in a liquid.

Homogeneous nature of a solution

When a few crystals of potassium permanganate are added to water, a purple coloured layer is seen at the bottom of the beaker covering the crystals. After sometime, the colour spreads throughout the water. The intensity of the colour is however, not uniform. On stirring the solution well, the crystals dissolve completely. The colour of the solution also becomes uniform. From this we may conclude that the crystals of potassium permanganate are broken into very small particles and are distributed uniformly or homogeneously throughout the solution. Thus *a solution is a homogeneous mixture of two or more substances*

In a solution, the solid that is dissolved is usually called the **solute** and the liquid is called the **solvent**. Water is a good solvent for many solid substances. Solutions prepared by using water as solvent are called **aqueous solutions** ('aqua' in Latin means 'water').

Heat changes accompanying dissolution of substances in water

When ammonium chloride is dissolved in water, the solution formed is cooler than water. It means that when the solid goes into solution, heat is absorbed Such a process in which heat is absorbed is called an **endothermic** process. But when sodium hydroxide is dissolved in water, a solution warmer than water is formed. This indicates that heat is liberated when sodium hydroxide dissolves in water. A process such as this in which heat is evolved or liberated is called an **exothermic** process. In the case of the dissolution of some substances like sodium chloride the heat change is too small to be noticed.

Can we recover the solute from a solution ?

When common salt is added to water and stirred, the salt disolvess and forms a solution. How can the salt be recovered ? The solution is taken in an evaporating dish and gently heated till all the water is driven off The residue left behind is salt. We know that common salt is obtained by evaporating sea water in salt pans.

Thus evaporation enables us to recover the solute from a solution. The solvent however, escapes into the atmosphere in the process of evaporation. To recover the solvent from the solution, we make use of a distillation apparatus as shown in Figure 4.1.

A solution of common salt in water is taken in a distillation flask attached to a condenser and a receiver. The flask is heated. The water (solvent) evaporates and passes through the condenser. The water vapour condenses to a liquid as it passes through the condenser which is kept cool by circulating water through it. the condensed liquid (distillate) is collected in the receiver.

It may be mentioned here that the distilled water used in the laboratory and in car batteries is a pure sample of water obtained by distilling water.

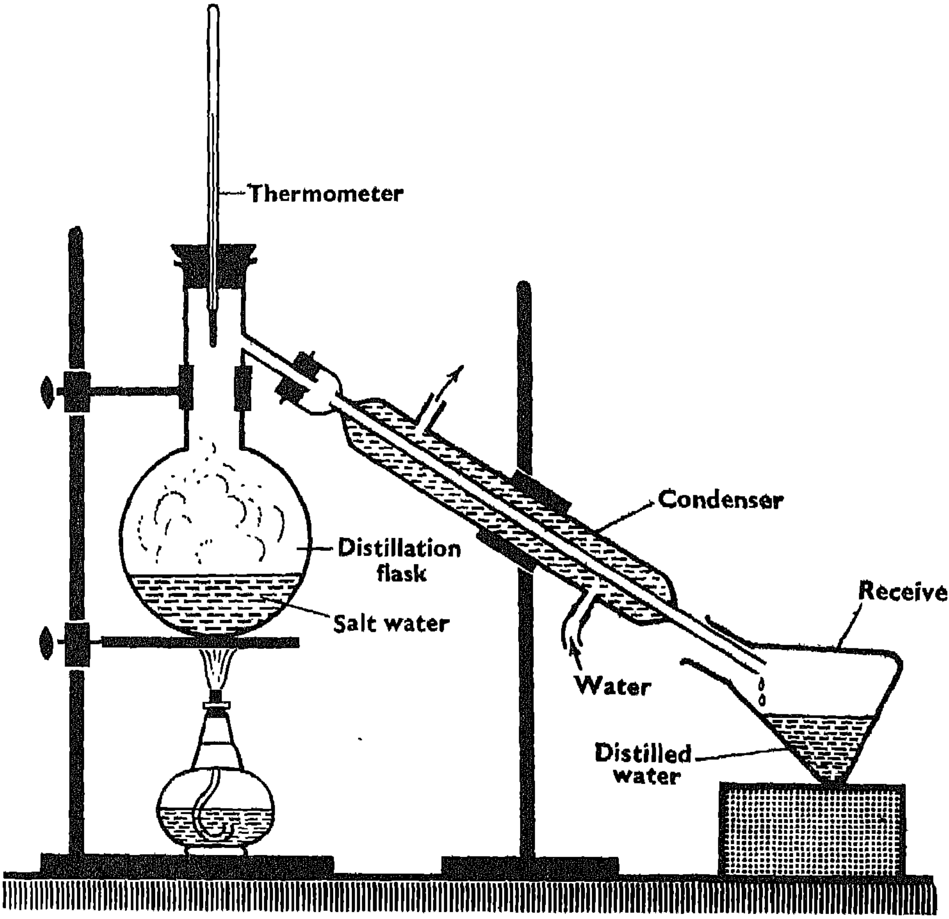


Figure 4.1 *Distillation apparatus*

4.2 Process of dissolution.

When a substance like salt or sugar is placed in water, the substance is found to disappear gradually into the solvent. That is, a solution is formed. During this process, the forces of attraction holding the solute particles together are somehow weakened. Stronger forces of attraction between the solute and solvent particles come into play. As a result, particles of the solute drift into the solvent and form a solution (Figures 4.2 and 4.3).

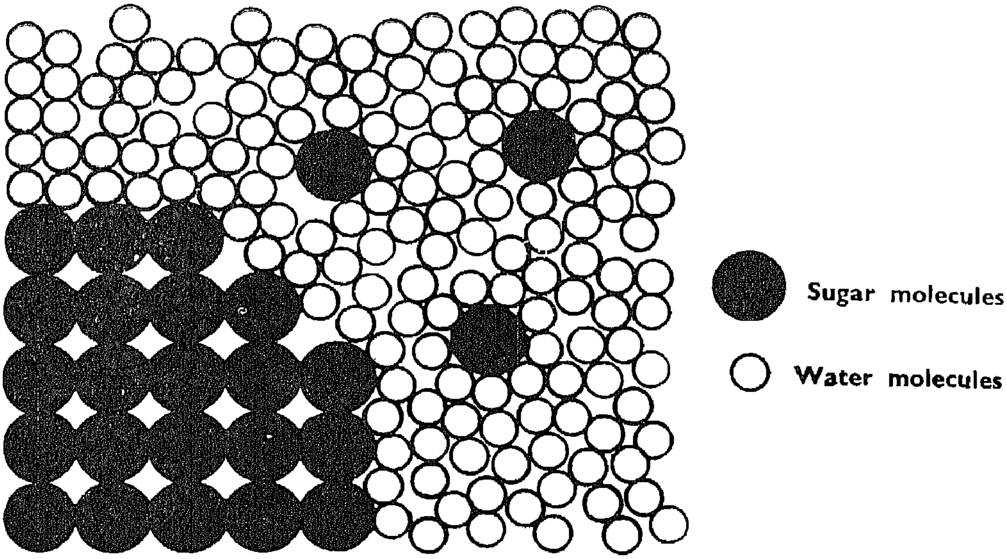


Figure 4.2 *Dissolution of sugar in water*

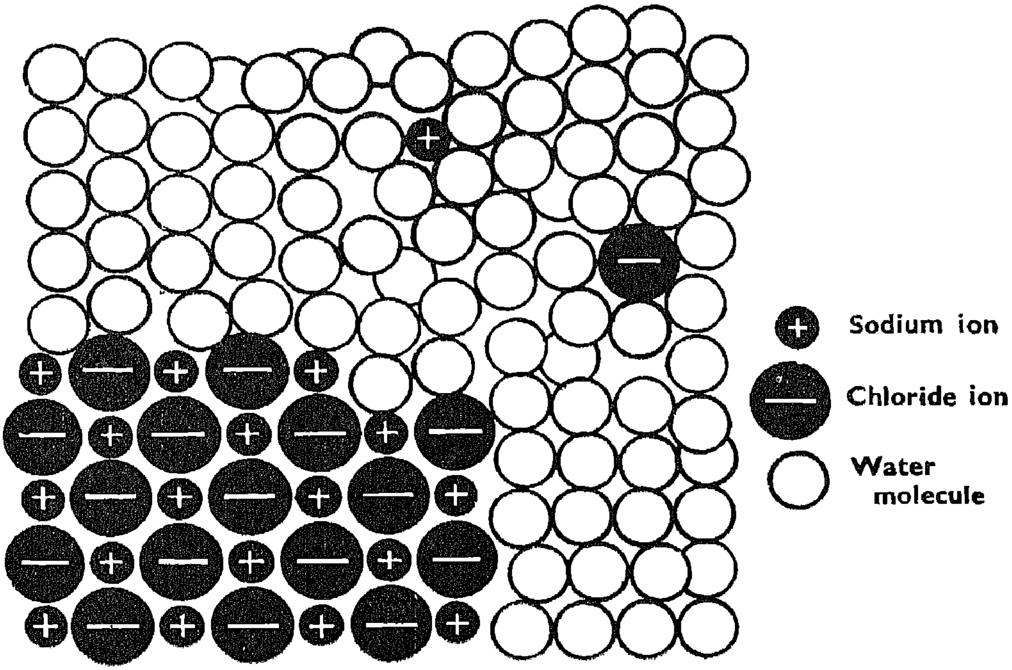


Figure 4.3 *Dissolution of sodium chloride in water*

We find from experiment that a powdered solid dissolves faster than a lump. Stirring and heating also speed up dissolution.

4.3 Saturated and unsaturated solutions

Let us take a small quantity of water and add common salt to it in small portions. We find that the salt dissolves readily at first, but soon a stage is reached when no more salt goes into solution. At this stage the added salt remains undissolved in the solution. Thus only a limited amount of salt dissolves in a given volume of water at a particular temperature A solution that does not dissolve any more solute at a given temperature is called a **saturated solution** at that temperature. A solution which contains less amount of solute than required for saturation is called an **unsaturated solution**.

Crystals and crystallisation

On cooling a hot saturated solution of potassium nitrate, the salt separates out in the form of crystals. The process of obtaining a substance in the form of crystals is called **crystallisation**. By slow evaporation of a saturated solution of a substance, it is possible to prepare crystals of large size. Crystals of copper sulphate are obtained by dissolving crude copper sulphate in water and concentrating the solution.

When examined under a magnifying glass, crystals are found to have definite geometrical shapes, with flat regular surfaces. A crystal of common salt is cubic in shape.

Concentration of solutions

The amount of a solute dissolved in a specified quantity of a solution is a measure of the concentration of a solution. If the amount of solute is small, we get a dilute solution. If it is large, we get a concentrated solution. It may be noted that 'dilute' and 'concentrated' are only relative terms. They do not denote any specific quantity of the solute dissolved.

100 ml of a solution containing 2 g of sodium chloride is called a 2 per cent solution of sodium chloride. The concentration of a solution is thus commonly expressed as the number of grams of solute in 100 ml of the solution.

4.4 Electrical conductivity of aqueous solutions

It is commonly known that an electric current passes readily through a copper wire, but does not pass through a piece of wood. Copper, therefore, is a conductor and wood a non-conductor.

It has been shown that water is not a good conductor of electricity, but a solution of copper sulphate in water readily conducts electricity. A solution of sugar in water, on the other hand, does not conduct electricity (Fig. 4.4).

As its aqueous solution readily conducts electricity, copper sulphate is called an **electrolyte.** Sugar, whose aqueous solution does not conduct electricity, is called a **non-electrolyte.**

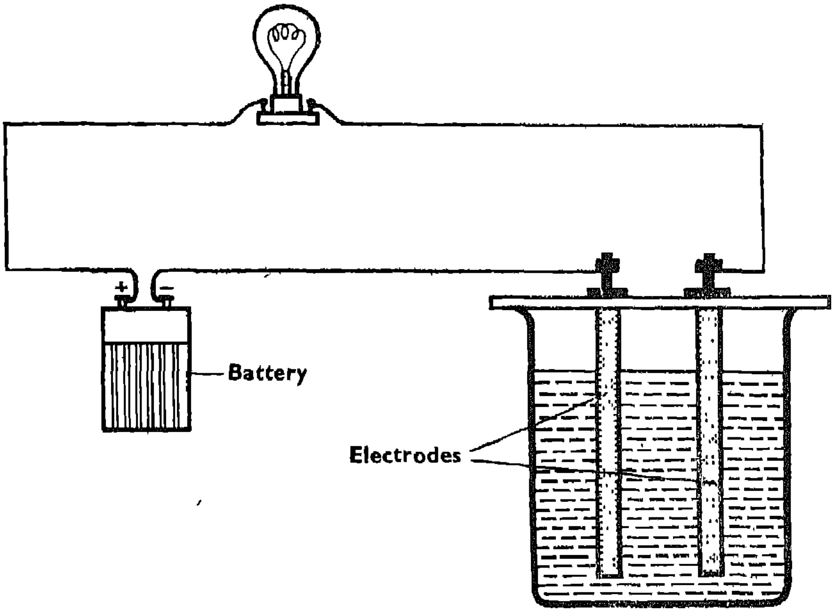


Figure 4.4 *Electrical conductivity of aqueous solutions*

An aqueous solution of copper sulphate contains cupric ions, Cu^{++} and sulphate ions, SO_4^{--}. These ions, being charged particles, move towards the oppositely charged electrodes when an electric field is applied. This movement of ions makes the solution conduct electricity. On the other hand, a solution of sugar in water contains no ions but only neutral molecules. Such a solution does not conduct electricity.

4.5 Types of solutions

The solute and solvent may be in one of the three states of matter—solid, liquid or gaseous. It is thus possible to have different types of solutions.

Solids in liquids

So far we have examined only solutions of solids in liquids, such as common salt in water. Such solid-liquid solutions are the ones commonly met with in everyday life and in the laboratory. The liquid generally used is water.

Solids which do not dissolve in water, may dissolve in other solvents. Paraffin wax which is insoluble in water is soluble in kerosene. Iodine dissolves only to a small extent in water, but readily dissolves in alcohol.

Liquids in liquids

Alcohol mixes with water in all proportions. A solution of alcohol in water is thus obtained. Kerosene, however, does not mix with water. It is said to be immiscible with water.

Gases in liquids

When a bottle of soda water is opened, gas bubbles are seen coming out. The gas evolved is carbon dioxide. It dissolves in water to some extent, forming a solution.

In the preparation of soda water, carbon dioxide gas at a high pressure is passed into water and then corked. When

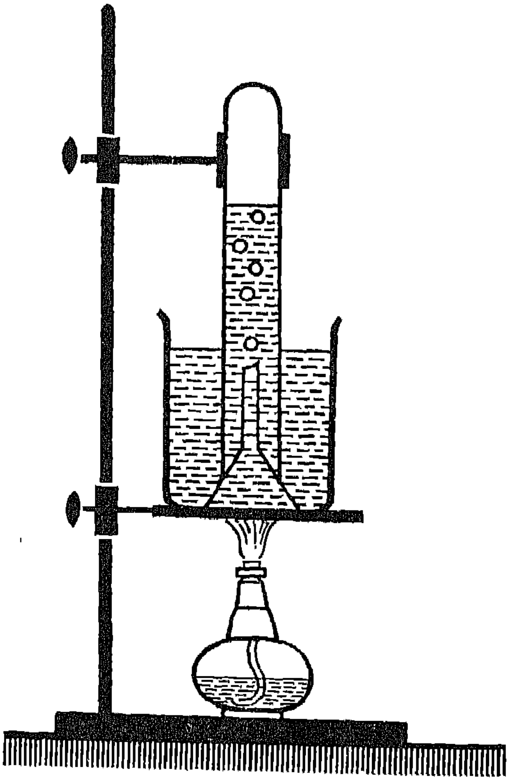


Figure 4.5 *Solubility of air in water*

the cork is opened, the pressure is released and so the excess carbon dioxide bubbles out of the solution

Is air soluble in water ? It is observed that when a beaker containing water is heated, bubbles of air come out of water, and collect in the test tube (Fig. 4.5).

This shows that air is soluble in cold water, but less to in hot water. Generally the solubility of gases in liquids decreases with increasing temperature.

Some gases like hydrogen chloride and ammonia are highly soluble in water. This is shown by the fountain experiment (Fig. 4.6).

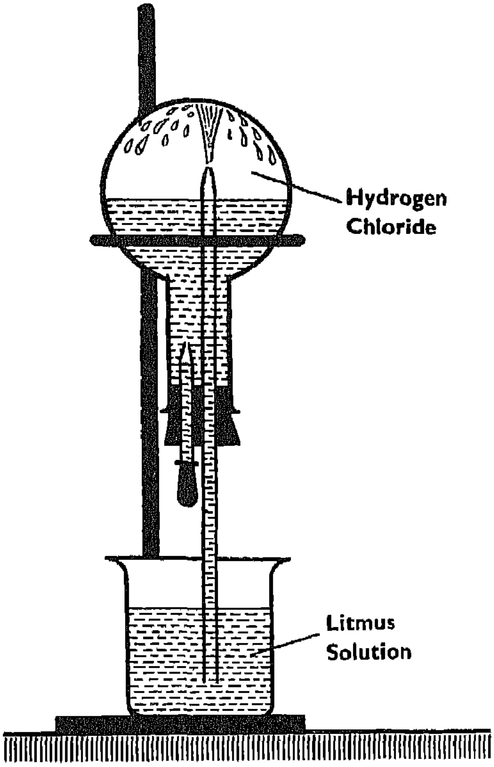


Figure 4.6 *Solubility of hydrogen chloride in water*

Hydrogen chloride gas is collected in a dry round bottomed flask. The flask is fitted with a two holed rubber stopper. A dropper containing a small amount of water passes through one hole and a long tube with a nozzle, through the other. The flask is inverted in such a way that the glass tube dips in blue litmus solution taken in a beaker. A few drops of water are introduced into the flask by squeezing the dropper. Blue litmus solution from the beaker rushes into the flask in the form of a fountain. The colour of the solution turns red.

The hydrogen chloride gas dissolves in the water introduced into the flask. A partial vacuum is thereby created. To fill the vacuum the blue litmus solution rushes into the flask. The hydrogen chloride solution turns blue litmus red.

Role of solutions in daily life

We come across different types of solutions in our daily life. Sea water is a solution of sodium chloride and other salts in water. Tincture of iodine (used as an antiseptic) is a solution of iodine and potassium iodide in alcohol. Soda water is a solution of carbon dioxide in water. Chlorine gas is sometimes dissolved in water to kill bacteria. Oxygen dissolved in water supports marine life. Fertilizers like ammonium sulphate, when added to soil dissolve in water and are assimilated by plants.

Solutions thus play a vital role in many processes.

Questions

1. What is meant by solute, solvent and solution ?
2. What is an aqueous solution ? How is an aqueous solution of copper sulphate prepared ?
3. What are exothermic and endothermic processes ?
4. Is the dissolution of sodium hydroxide in water exothermic or endothermic ?
5. After adding sugar to coffee, we generally stir the liquid with a spoon. Why ?
6. What are the factors that quicken the process of dissolution of a salt in water ?
7. What is meant by a saturated solution ? Describe how a saturated solution of potassium dichromate is prepared at room temperature.
8. What is the difference between saturated and unsaturated solutions ? How will you determine whether a given solution is saturated or not ?

9. How do you prepare pure crystals of copper sulphate from a commercial sample ?

10. A salt solution conducts electricity but sugar solution does not. Explain.

11. Suggest suitable solvents for naphthalene, paraffin wax and iodine.

12. Explain why gas bubbles out when a soda water bottle is opened?

13 How can the presence of air dissolved in water be demonstrated ?

14. Choose the correct word to complete the following statements
    a. Dissolution of ammonium chloride in water is an . ... process (exothermic/endothermic)
    b. The solubility of hydrogen chloride in water is .. . . (high/low).
    c. A substance, an aqueous solution of which conducts electricity, is called.. . . ..(electrolyte/non-electrolyte)
    d The solubility of gases generally . ... with increasing temperature (increases/decreases)
    e Stirring. . ...the rate of dissolution of a solute in a solvent (increases/decreases)

15. State whether the following statements are TRUE or FALSE.
    a. A solution is a homogeneous system.
    b. Powdered sugar dissolves faster than sugar candy
    c. When a salt is dissolved in water heat is always evolved.
    d. A solution of cane sugar in water conducts electricity.
    e. An unsaturated solution can dissolve more of the solute.
    f Air is more soluble in cold water than in hot water.
    g. Iodine dissolves readily in water but not in alcohol.
    h. Soda water is a solution of oxygen in water

UNIT 5

Chemical Reactions

The study of chemical reactions is an important aspect of chemistry. In this unit we shall learn what a chemical reaction is and what its characteristics are. We shall also study the conditions for its occurrence and the factors affecting its rate. The more common types of reactions are also discussed.

5.1 What is a chemical reaction ?

In unit 1 we studied a number of chemical changes such as burning of magnesium in air, electrolysis of water, rusting of iron and burning of wood. In such changes new substances are formed. It is observed that when copper carbonate is heated, it gives black cupric oxide and carbon dioxide.

$$\underset{\text{Copper carbonate}}{CuCO_3} \longrightarrow \underset{\text{Cupric oxide}}{CuO} + \underset{\text{Carbon dioxide}}{CO_2\downarrow}$$

Similarly, mercuric oxide, on heating, gives shining globules of mercury and oxygen gas which rekindles a glowing splinter.

$$\underset{\text{Mercuric oxide}}{2\,HgO} \longrightarrow \underset{\text{Mercury}}{2Hg} + \underset{\text{Oxygen}}{O_2\uparrow}$$

On placing an iron nail in a solution of copper sulphate, the nail is coated with a reddish-brown substance, copper. The intensity of the blue colour of the solution fades.

$$\underset{\text{Iron}}{Fe} + \underset{\text{Copper sulphate}}{CuSO_4} \longrightarrow \underset{\text{Ferrous sulphate}}{FeSO_4} + \underset{\text{Copper}}{Cu\downarrow}$$

In all such changes it is noticed that the new substances formed have properties different from those of the reactants. *Such changes are known as* **chemical reactions**.

A large number and a variety of reactions take place in nature. Many reactions are also carried out in laboratories and factories. Materials of everyday use, whether obtained in nature (e.g., water, common salt, petroleum etc) or from a factory (e.g., soap, paper, drugs etc) are products of chemical reactions

5.2 Characteristics of a chemical reaction

When water is added to quicklime (calcium oxide) much heat is liberated. This may have been observed during the preparation of slaked lime (calcium hydroxide) used for white washing

$$\underset{\text{Calcium oxide}}{CaO} + \underset{\text{Water}}{H_2O} \longrightarrow \underset{\text{Calcium hydroxide}}{Ca(OH)_2}$$

When magnesium burns in air, it produces a dazzling light and also much heat.

$$\underset{\text{Magnesium}}{2\,Mg} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Magnesium oxide}}{2\,MgO}$$

When a magnesium ribbon and a copper strip dipped in dilute sulphuric acid are connected externally to a torch bulb through metallic wires, the bulb glows. The glowing of the bulb is due to electrical energy liberated as a result of chemical reaction. Thus in chemical reactions energy is released—usually in the form of heat and sometimes as light or electrical energy. However in some reactions energy is also consumed, as for example, in the electrolysis of water All chemical reactions are therefore accompanied by evolution or absorption of energy. In some reactions the energy changes may be too small to be observed, as in the case of the rusting of iron.

5.3 Conditions that help bring about a chemical reaction

Let us consider under what conditions a chemical reaction occurs.

i) Intimate contact

When silvery white mercury and yellow sulphur powder are placed together in a mortar, practically no reaction takes place. But on grinding the two together, a black powder (mercuric sulphide) is formed. Grinding helps in providing an intimate contact between the reactants. Intimate contact between reactants is thus a primary condition for a reaction to occur.

ii) Energy

Does intimate contact alone bring about chemical reaction in all cases ? Let us study some more reactions.

A mixture of powdered copper and sulphur undergoes no appreciable change on grinding or on keeping for a long time. But on heating the mixture, reaction sets in and black copper sulphide is formed.

$$\underset{\text{Copper}}{Cu} + \underset{\text{Sulphur}}{S} \longrightarrow \underset{\text{Copper sulphide}}{CuS}$$

It is evident from the above example, that certain chemical reactions take place when heat energy is supplied. Another reaction in which heat is to be supplied is the formation of iron sulphide from iron and sulphur.

It may be recalled that plants prepare starch in the presence of sunlight (light energy) by a process known as photosynthesis When silver chloride (white), or silver bromide (pale yellow), is exposed to bright sunlight, the colour changes to grey or black. But no such change occurs when these substances are kept in the dark. The change in colour is due to a partial decomposition of the substances under the action of light. To prevent such decomposition, silver compounds are stored in dark brown bottles ; thereby much of the light is cut off . The action of light on silver bromide is made use of in photography.

The examples mentioned above show that some chemical reactions take place by the absorption of light energy.

It has already been mentioned that when an electric current is passed through water containing a small quantity of common salt, oxygen and hydrogen are liberated.

$$\underset{\text{Water}}{2\,H_2O} \xrightarrow{\text{Electrolysis}} \underset{\text{Hydrogen}}{2\,H_2\uparrow} + \underset{\text{Oxygen}}{O_2\uparrow}$$

We are familiar with articles made of brass, such as spoons, tumblers, etc., which are coated with a metal like nickel. This coating is done by dipping the article in a solution containing some suitable compound of nickel. The article to be coated is connected to the negative terminal of a battery and a nickel plate, dipped in the solution, is connected to the positive terminal. On passing the current, electrolysis takes place and nickel is deposited on the article. In this process, known as electro-plating, a chemical reaction brought about by electrical energy, makes it possible for nickel to be deposited on the article.

Electroplating is widely employed in coating a number of articles such as automobile parts, bicycle parts, surgical instruments, etc. with a suitable metal.

A mixture of sulphur (1 part) and potassium chlorate (2½ parts) does not undergo any reaction if kept undisturbed. But if it is wrapped in a paper as in a cracker and struck with a hammer, a loud explosion results. Similarly a cartridge may be left in the barrel of a gun for years without any change. It explodes only when it is hit by the trigger. Certain reactions therefore, occur when reactants are suddenly struck (percussion).

Thus, in addition to intimate contact, energy in some form or other may be needed to bring about a reaction.

5.4 Factors which influence the rate of a reaction

a. Particle size

When zinc granules react with dilute hydrochloric acid there is a slow evolution of hydrogen. On the other hand, when zinc dust reacts with the acid, there is a brisk evolution of the gas. Similarly finely powdered marble (calcium carbonate) reacts more vigorously with dilute hydrochloric acid (to liberate carbon dioxide gas) than a lump of marble. Particle size, thus, has a considerable influence upon the rate of a reaction. The smaller the particles, the more rapid is the reaction. In the case of solid reactants, the reaction takes place at the surface of contact of the solids. When solids are finely divided, the surface area is greatly increased. This increase in the surface permits a greater contact between reactants. Hence, the chances of reaction are greater than in the coarse state.

b Concentration

When equal volumes of concentrated and dilute hydrochloric acid are added to two test tubes containing equal amounts of zinc granules (commercial), evolution of hydrogen gas is more brisk with the concentrated acid than with the dilute. Concentration, therefore, is another factor influencing the rate of a reaction. Larger the concentration, the faster is the reaction.

c Temperature

When a solution of oxalic acid containing dilute sulphuric acid is added to a solution of potassium permanganate, the purple colour of the permanganate disappears slowly. But if a hot solution of oxalic acid is added, the decolourisation takes place more rapidly. Similarly evolution of hydrogen gas is more brisk when zinc is added to hot hydrochloric acid than to the cold acid. Temperature, therefore, influences the rate of a reaction. In general, reactions occur faster at a higher temperature than at a lower temperature.

d. Catalyst

As mentioned earlier, if a solution of oxalic acid is added to an acidified solution of potassium permanganate, the purple colour of the permanganate disappears very slowly. However, if a small amount of manganese sulphate is added to the solution, the purple colour disappears rapidly. Manganese sulphate remains unchanged at the end of the reaction. It merely alters the speed of a reaction.

Such substances which alter the speed of a reaction, but remain unchanged at the end of the reaction are called **catalysts**.

Thus particle size, concentration, temperature and catalysts are some of the factors which influence the rate of a reaction.

6.5 Types of chemical reactions

Following are the common types of reactions.

(i) Magnesium burns in air forming magnesium oxide.

$$\underset{\text{Magnesium}}{2\,Mg} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Magnesium oxide}}{2\,MgO}$$

Ammonia gas and hydrogen chloride react to give ammonium chloride.

NH_3	+	HCl	$\longrightarrow$	NH_4Cl
Ammonia		Hydrogen chloride		Ammonium chloride

Reactions such as these, *in which two substances, whether elements or compounds, combine to give a single new substance are called* **combination** *reactions.*

(ii) On heating, mercuric oxide breaks down (decomposes) into its elements, mercury and oxygen.

$2\,HgO$	$\longrightarrow$	$2\,Hg$	+	$O_2\uparrow$
Mercuric oxide		Mercury		Oxygen

Similarly, on heating, copper carbonate decomposes into cupric oxide and carbon dioxide.

$CuCO_3$	$\longrightarrow$	CuO	+	$CO_2\uparrow$
Copper carbonate		Cupric oxide		Carbon dioxide

Reactions such as these, *in which a compound breaks down into two or more simpler substances, either elements or compounds, are called* **decomposition** *reactions.*

(iii) When an iron strip is placed in a solution of copper sulphate, iron dissolves gradually and copper is deposited.

Fe	+	$CuSO_4$	$\longrightarrow$	$FeSO_4$	+	$Cu\downarrow$
Iron		Copper sulphate		Ferrous sulphate		Copper

Iron displaces copper from copper sulphate solution. *A reaction such as this, in which one element displaces another in a compound is called a* **displacement** *reaction.*

(iv) When solutions of potassium iodide and lead nitrate are mixed, lead iodide is precipitated. Potassium nitrate remains in solution.

$2KI$	+	$Pb(NO_3)_2$	$\longrightarrow$	$2KNO_3$	+	$PbI_2\downarrow$
Potassium iodide		Lead nitrate		Potassium nitrate		Lead iodide

$$2K^+ + 2I^- + Pb^{++} + 2NO_3^- \longrightarrow 2K^+ + 2NO_3^- + PbI_2\downarrow$$

Similarly, when dilute sulphuric acid is added to a solution of barium chloride, barium sulphate precipitates out. Hydrochloric acid remains in solution.

$$\underset{\text{Barium chloride}}{BaCl_2} + \underset{\text{sulphuric acid}}{H_2SO_4} \longrightarrow \underset{\text{barium sulphate}}{BaSO_4\downarrow} + \underset{\text{hydrochloric acid}}{2HCl}$$

$$Ba^{++} + 2Cl^{-} + 2H^{+} + SO_4^{--} \longrightarrow BaSO_4\downarrow + 2H^{+} + 2Cl^{-}$$

Such reactions, which involve a mutual exhange of ions are known as **double decomposition** *reactions.*

It is important to note that in every chemical reaction, the substances formed (*products*) have different properties from those of the starting substances (*reactants*).

Questions

1. List some examples of chemical reactions which you have studied. Represent these reactions by equations
2. Give examples to show that (a) heat energy, (b) light energy and (c) electrical energy help bring about chemical reactions.
3. State whether the following statements are TRUE or FALSE.
    - a. The larger the particle size of the reactants, the faster is the reaction.
    - b. The lower the concentration the faster is the reaction.
    - c. The speed of a reaction is higher at a higher temperature than at a lower temperature.
    - d. A catalyst alters the rate of a chemical reaction.
    - e. The reaction between magnesium and oxygen to give magnesium oxide is a decomposition reaction.
    - f. Chemical reactions are always accompanied by the evolution of energy.
    - g. Intimate contact alone may not always bring about certain chemical reactions
    - h. In the electrolysis of water electrical energy is produced.
    - i The principle of photography involves the action of light on silver bromide.
4. State whether energy is released or absorbed in the following reactions:
    - a. Direct current is passed through water containing a small amount of salt.
    - b. Magnesium ribbon burns in air.
    - c. Magnesium ribbon and copper plates dipped in dilute sulphuric acid are connected externally to a torch bulb by means of connecting wires.

d. Charcoal burns.
e. A photograph is taken in good day light.

5. Mark in which of the following processes a chemical reaction has taken place.
a. A piece of gold is drawn into a thin wire.
b. Quicklime is added to water.
c. A piece of iron is exposed to moist air

6. Explain why
a. burning of magnesium is a chemical reaction.
b. evolution of carbon dioxide in the reaction between marble and hydrochloric acid is faster when marble is in the form of powder than in the form of lump.
c. silver bromide turns grey or black on exposure to light.
d. the colour of copper sulphate solution fades when iron nails are placed in it.

7. Choose the proper agent from column 'B' for bringing about the reactions mentioned in column 'A'.

'A'	'B'
a. Decomposition of water	Light
b. Formation of starch in plants	Electricity
c. Decomposition of mercuric oxide	Heat
	Catalyst

8. Fill up the blanks.
a. A change which leads to the formation of new substances is called a ... .
b. The . ...particle size, the faster is the reaction.
c. The reaction between hydrogen and oxygen to form water is a . .. .reaction.
d. Iron...... ....copper from copper sulphate solution
e. Mercuric oxide, on heating,.. ... .to mercury and oxygen.
f. Heat is... .. ... .when quicklime reacts with water to form slaked lime.
g. A catalyst ....the speed of a reaction.

9. Tick off the correct answer in the following cases.
a. The colour of an acidified solution of potassium permanganate is found to disappear slowly on addition of oxalic acid. The colour can be made to disappear more rapidly by
   (i) heating the solution
   (ii) cooling the solution
   iii) exposing the solution to light

b. Reaction of iron with copper sulphate solution is an example of

(i) decomposition

(ii) displacement

(iii) double decomposition

(iv) combination

10 What type of reaction does each of the following equations indicates ?

a. $2\,NaCl + Pb(NO_3)_2 \longrightarrow PbCl_2\downarrow + 2\,NaNO_3$

b. $Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2\uparrow$

c. $CaCO_3 \longrightarrow CaO + CO_2\uparrow$

d. $C + O_2 \longrightarrow CO_2\uparrow$

e. $Cu + 2\,AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag\downarrow$

11. A solution containing 5 g of a substance 'A' dissolved in 100 ml of water is added to a solution containing 8 g of substance 'B' in 100 ml water. Another solution containing 9 g of 'A' dissolved in 200 ml of water is added to a solution of 'B' of the same concentration as above. In which case is the reaction between 'A' and 'B' faster? Explain.

UNIT 6

Some Familiar Substances

The air we breathe contains oxygen, nitrogen and small amounts of carbon dioxide. The water we drink is a compound of hydrogen and oxygen. Common salt, which is used in our food, is a compound of sodium and chlorine. All food materials and fuels contain carbon. Matches, some of the medicines and crackers contain sulphur. In this unit we shall study these familiar substances and some of their constituents.

6.1 Air

We learnt in unit 1 about some of the common substances such as air, water, charcoal, sulphur and salt.

Air is a mixture mainly of oxygen and nitrogen. These constitute about one-fifth and four-fifths of the volume of air respectively. When air is bubbled through lime water, the latter turns milky. This shows the presence of carbon dioxide in air.

Early in the mornings during winter you might have observed droplets of water on grass, stones and sand, even though there had been no rain. Water droplets form on the outer surface of a beaker containing ice. These observations indicate the presence of water vapour in air.

Thus, air contains, besides oxygen and nitrogen, small quantities of carbon dioxide and water vapour. Although the composition of air is nearly constant it varies slightly from place to place.

William Ramsay and others have shown that air contains in addition to the above, small proportions (less than 1%) of rare gases —argon, neon, helium etc.

Oxygen and nitrogen are essential to plant and animal life. Carbon dioxide is used by plants to make starch. We shall study here about oxygen, nitrogen, and carbon dioxide.

6.2 Oxygen

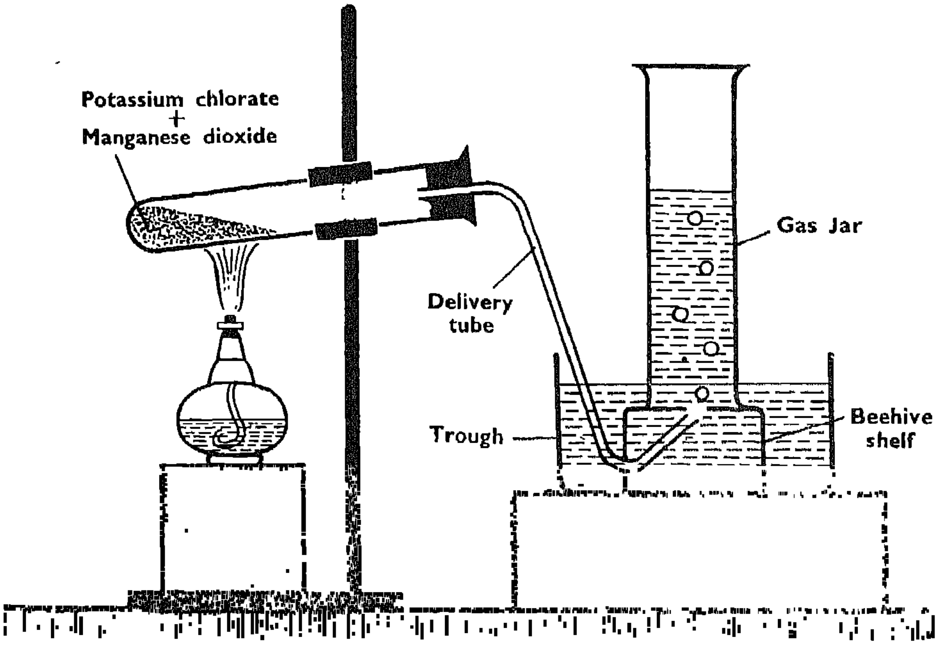
Without oxygen life is impossible even for a few minutes. Besides being present in air and water, oxygen occurs abundantly in the earth's crust as sand (silicon dioxide) and as oxides of many other elements.

Preparation

In the laboratory, oxygen is usually prepared by heating potassium chlorate.

$$\underset{\text{Potassium chlorate}}{2\,KClO_3} \longrightarrow \underset{\text{Potassium chloride}}{2\,KCl} + \underset{\text{Oxygen}}{3O_2\uparrow}$$

Evolution of oxygen is slow if potassuim chlorate is heated alone. If, however, a mixture of potassium chlorate and a little manganese



Figures 6.1 *Preparation of oxygen*

dioxide (MnO_2) is heated, oxygen is evolved more readily. Here manganese dioxide acts as a catalyst. It increases the rate of evolution of oxygen. Oxygen formed is collected over water (Fig. 6.1).

Oxygen may also be prepared by heating potassium permanganate.

$2\,KMnO_4$	$\longrightarrow$	K_2MnO_4	+	MnO_2	+	$O_2 \uparrow$
Potassium permanganate		Potassium manganate		Manganese dioxide		Oxygen

Properties

Oxygen is a colourless and odourless gas. It is slightly soluble in water. When a burning splinter is introduced into oxygen, the gas does not burn but the splinter burns more brightly. A glowing splinter bursts into flame in oxygen. A piece of glowing charcoal burns brightly forming carbon dioxide which turns lime water milky.

C	+	O_2	$\longrightarrow$	$CO_2 \uparrow$
Carbon		Oxygen		Carbon dioxide

Sulphur burns in oxygen with a bright blue flame forming sulphur dioxide, a colourless gas with a suffocating smell.

S	+	O_2	$\longrightarrow$	$SO_2 \uparrow$
Sulphur		Oxygen		Sulphur dioxide

Red hot copper wire burns forming a black substance, cupric oxide.

$2\,Cu$	+	O_2	$\longrightarrow$	$2\,CuO$
Copper		Oxygen		Cupric oxide

Magnesium burns with a dazzling white light forming magnesium oxide.

$2\,Mg$	+	O_2	$\longrightarrow$	$2\,MgO$
Magnesium		Oxygen		Magnesium oxide

Similarly, many other substances like phosphorus and iron burn in oxygen indicating that oxygen is a reactive element. It combines with most of the elements forming the corresponding oxides. Such a reaction is called oxidation. Oxygen required for oxidation may come from air or from substances such as potassium chlorate, potassium permanganate and others. Such substances are called oxidising agents.

Uses

Oxygen present in air is used for respiration by all living beings. Fishes and water plants, however, breathe oxygen dissolved in water. Oxygen is given to patients who have difficulty in breathing. In cases of suffocation caused by drowning or breathing of poisonous gases oxygen is given. Mountain climbers, sea-divers and cosmonauts use oxygen stored in cylinders (Fig. 6.2).

Figure 6.2 *A mountain climber carrying oxygen cylinder for respiration*

Oxygen is essential for burning processes. When a mixture of oxygen and acetylene gas* is burnt in a blow pipe or a torch much heat is liberated (the temperature of the flame is about 3000°C). Oxy-acetylene flame is used in welding steel and also for cutting holes in steel sheets (Fig. 6.3).

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*Acetylene is a gas containing hydrogen and carbon.

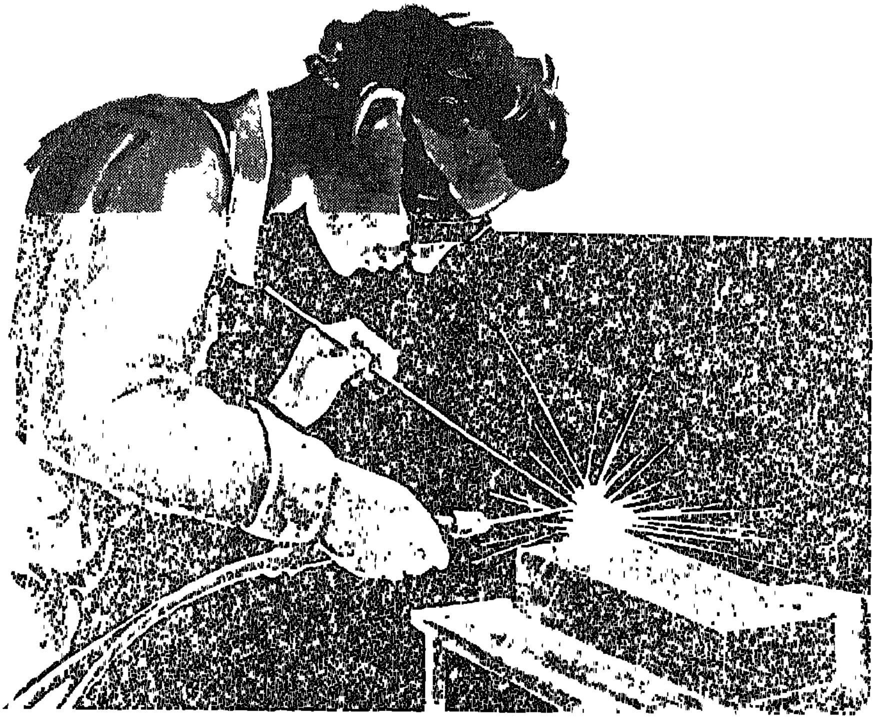


Figure 6.3 *Oxy-acetylene flame for welding*

6.3 Nitrogen

The nucleus of the living cell contains compounds of nitrogen. Nitrogen compounds form an essential constituent of the food of all living beings—men, animals and plants. Compounds of nitrogen are very widely used in the preparation of fertilizers, explosives, dyes and medicines.

Besides being present in air, nitrogen is also present as nitrates, e.g., saltpetre (potassium nitrate).

Preparation

Nitrogen is conveniently prepared in the laboratory by warming an aqueous solution of ammonium chloride and sodium nitrite. The gas is collected over water (Fig. 6.4).

$$\underset{\text{Ammonium chloride}}{NH_4Cl} + \underset{\text{Sodium nitrite}}{NaNO_2} \longrightarrow \underset{\text{Sodium chloride}}{NaCl} + \underset{\text{Water}}{H_2O} + \underset{\text{Nitrogen}}{N_2\uparrow}$$

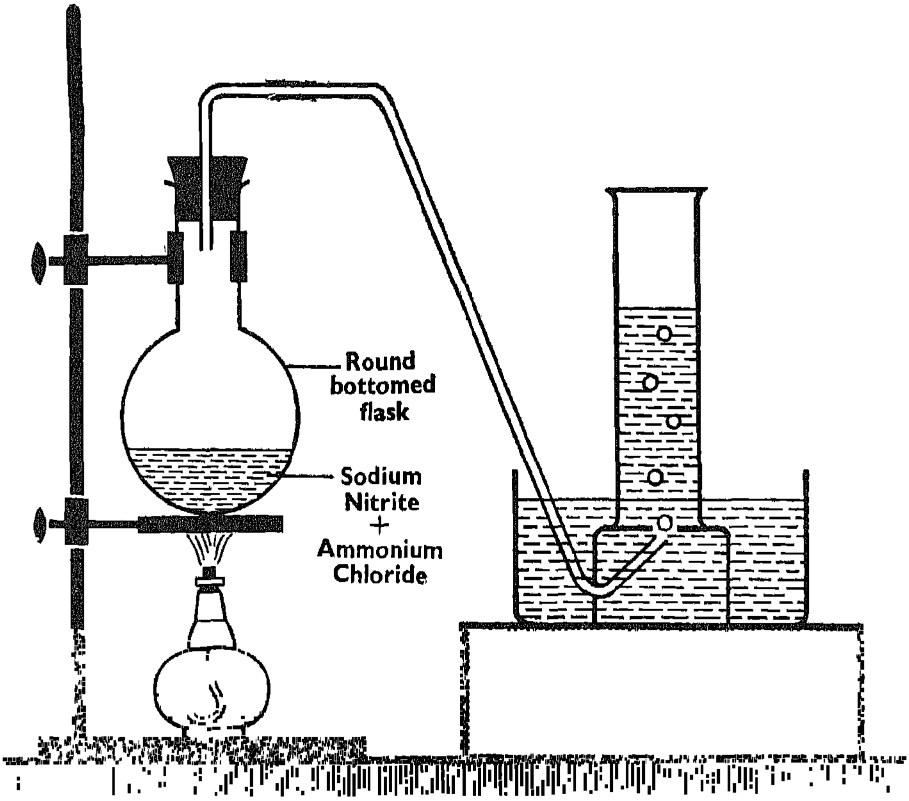


Figure 6.4 *Preparation of nitrogen*

Properties

Nitrogen is a colourless and odourless gas It is slightly soluble in water.

When a burning splinter is introduced into nitrogen, the splinter is put out and the gas does not burn.

Burning charcoal and sulphur are put out in nitrogen. Red hot copper does not burn. Burning magnesium ribbon, however, continues to burn with a white light forming magnesium nitride.

$$\underset{\text{Magnesium}}{3\,Mg} + \underset{\text{Nitrogen}}{N_2} \longrightarrow \underset{\text{Magnesium nitride}}{Mg_3N_2}$$

Under ordinary conditions, most substances do not burn in nitrogen indicating that it is not very reactive.

Uses

Nitrogen is essential for the growth of plants. Fertilizers such as ammonium sulphate, ammonium nitrate and urea contain nitrogen. Nitrogen is not essential for breathing and remains unchanged during the process. However, its presence in air dilutes the oxygen and slows down oxidation.

6.4 Carbon dioxide

Carbon dioxide is the product of respiration of all living beings. It is also produced during the burning and decay of organic matter. Plants synthesise their food from carbon dioxide of the air. Common beverages like soda-water contain carbon dioxide.

Preparation

In the laboratory, carbon dioxide is prepared by the action of dilute hydrochloric acid on marble (calcium carbonate). The gas is collected by the upward displacement of air (Fig. 6.5).

$$\underset{\substack{\text{Calcium}\\ \text{carbonate}}}{CaCO_3} + \underset{\substack{\text{Hydrochloric}\\ \text{acid}}}{2\,HCl} \longrightarrow \underset{\substack{\text{Calcium}\\ \text{chloride}}}{CaCl_2} + \underset{\text{Water}}{H_2O} + \underset{\substack{\text{Carbon}\\ \text{dioxide}}}{CO_2\uparrow}$$

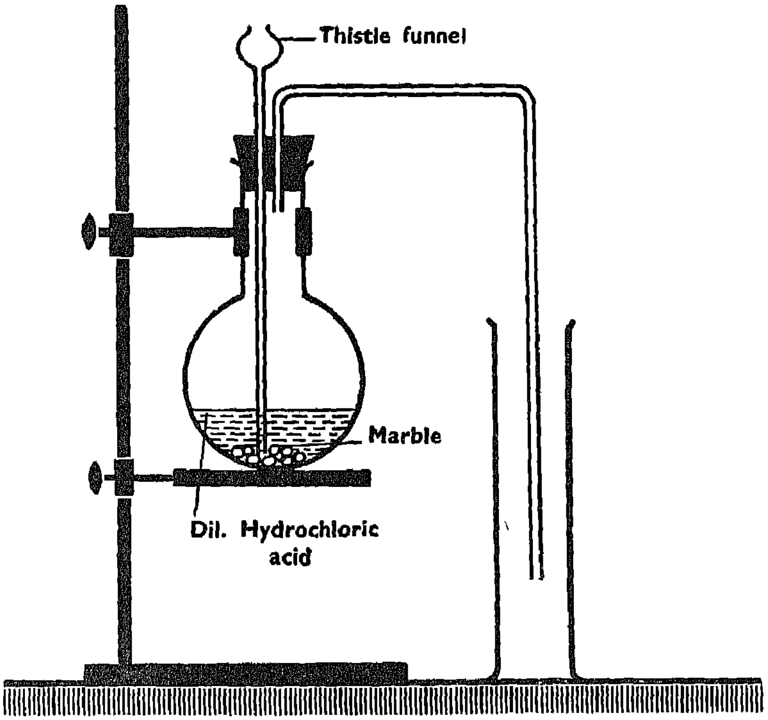


Figure 6.5 *Preparation of carbon dioxide*

Properties

Carbon dioxide is a colourless and odourless gas. It is heavier than air and fairly soluble in water.

When a burning splinter is introduced into carbon dioxide, the splinter is put out and the gas does not burn.

Burning sulphur is put out in carbon dioxide. Burning magnesium, however, continues to burn depositing carbon.

$$\begin{array}{ccccccc} 2\,Mg & + & CO_2 & \longrightarrow & 2\,MgO & + & C \\ \text{Magnesium} & & \text{Carbon dioxide} & & \text{Magnesium oxide} & & \text{Carbon} \end{array}$$

Lime water (calcium hydroxide) turns milky when carbon dioxide is bubbled through it.

$$\begin{array}{ccccccc} Ca(OH)_2 & + & CO_2 & \longrightarrow & CaCO_3\downarrow & + & H_2O \\ \text{Calcium hydroxide} & & \text{Carbon dioxide} & & \text{Calcium carbonate} & & \text{Water} \end{array}$$

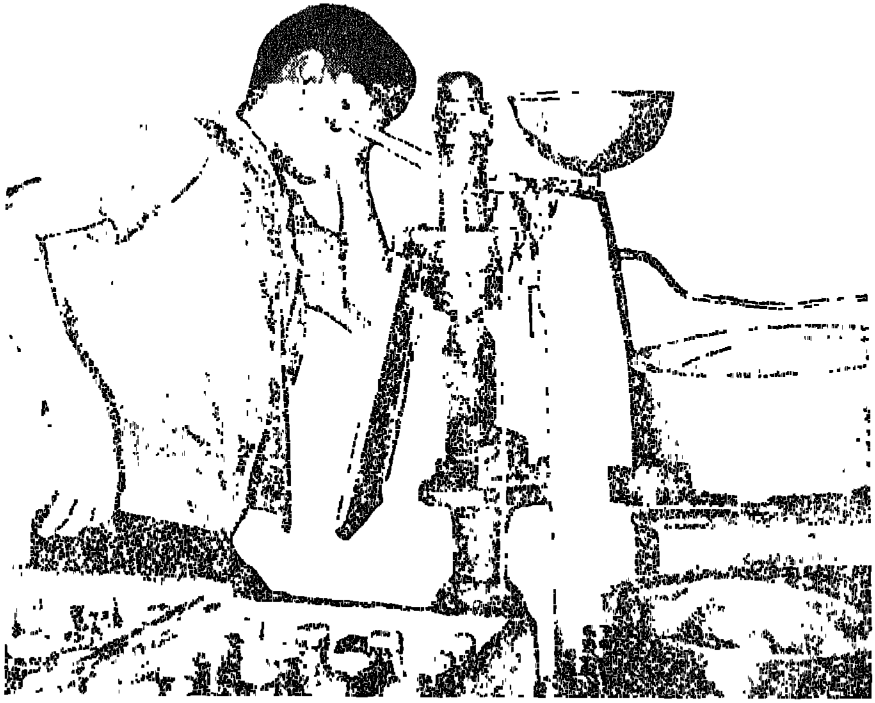


Figure 6.6 *Use of carbon dioxide in making aerated waters*

The milky appearance is due to the formation of calcium carbonate which is insoluble in water and so precipitates out of the solution.

Uses

The most common use of carbon dioxide is in the preparation of aerated waters (Fig. 6.6).

The gas can be easily solidified and the solid, called 'dry ice' finds use as a cooling agent The gas is also used in the preparation of washing soda (sodium carbonate) and baking soda (sodium bicarbonate).

Since carbon dioxide neither burns nor helps in the burning of substances, it is used as a fire extinguisher (Fig. 6.7). Being heavier than air, the gas covers the fire and prevents the access of oxygen.

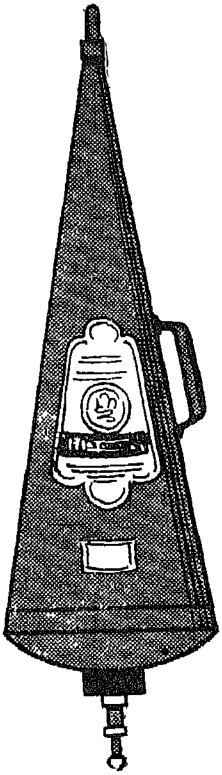


Figure 6.7 *A fire extinguisher*

6.5 Water

Next to air, the most important substance is water. About three-fourths of the surface of the earth is covered with water. It is the most abundant of all chemical compounds. The soil can retain water essential for the growth of plants. We can survive for several weeks without food but for only a few days without water. The digestion of food, the circulation of blood, the elimination of waste, the regulation of body temperature and many other functions depend upon an adequate supply of water. Thus water is essential to our existence

Water dissolves a large number of substances. It is, therefore, called a universal solvent.

On electrolysis, water gives hydrogen and oxygen.

$$\underset{\text{Water}}{2\,H_2O} \xrightarrow{\text{Electrolysis}} \underset{\text{Hydrogen}}{2\,H_2\uparrow} + \underset{\text{Oxygen}}{O_2\uparrow}$$

6.6 Hydrogen

Besides being a constituent of water, hydrogen is present in all living beings in combination with carbon, oxygen and nitrogen. Many familiar substances like petrol, sugar, oil etc., contain hydrogen.

Preparation

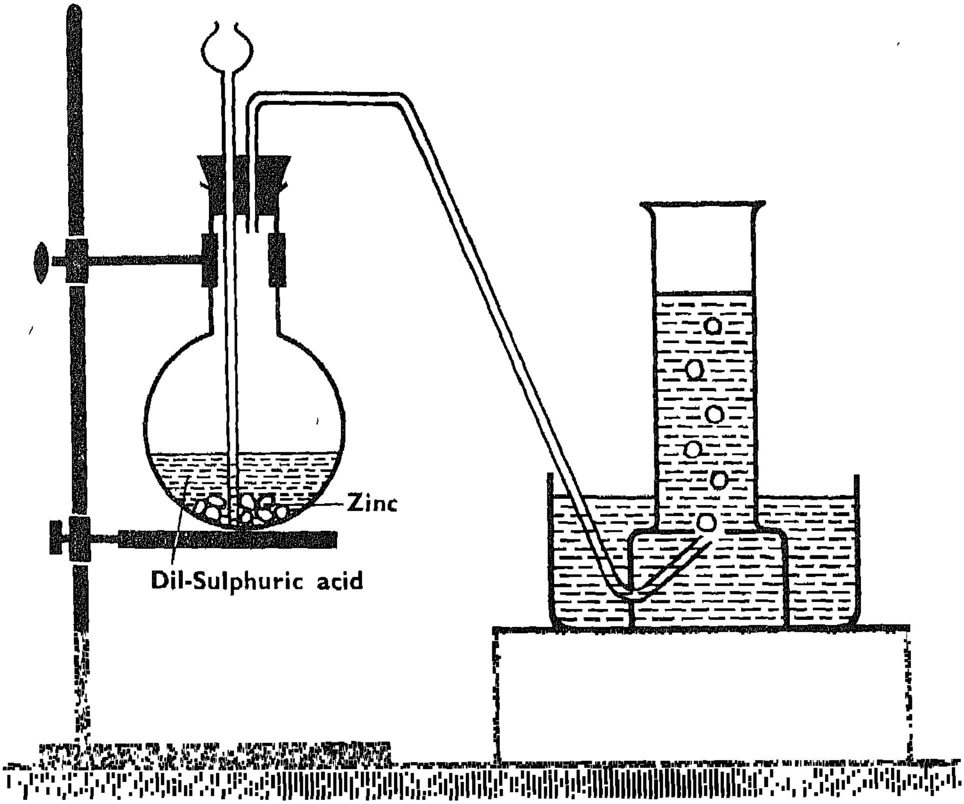
Hydrogen is prepared in the laboratory by the displacement of hydrogen present in acids, such as dilute hydrochloric or sulphuric acid, by metals like zinc or iron. Hydrogen formed is collected over water (Fig. 6.8).

$$\underset{\text{Zinc}}{Zn} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2} + \underset{\text{Hydrogen}}{H_2\uparrow}$$

$$\underset{\text{Zinc}}{Zn} + \underset{\text{Sulphuric acid}}{H_2SO_4} \longrightarrow \underset{\text{Zinc sulphate}}{ZnSO_4} + \underset{\text{Hydrogen}}{H_2\uparrow}$$

Properties

Hydrogen is a colourless and odourless gas. It is insoluble in water. A balloon filled with hydrogen when let loose, rises high into the air. If a jar of hydrogen is left open, the gas readily escapes into air. These observations show that hydrogen is lighter than air. In fact, it is the lightest substance known.



6.8 *Preparation of hydrogen*

Hydrogen burns in air forming water.

$$\underset{\text{Hydrogen}}{2\,H_2} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Water}}{2\,H_2O}$$

When a burning splinter is introduced into hydrogen it is put out with a 'pop' sound; the gas itself burns with a pale blue flame. The 'pop' sound is due to a mild explosion taking place when hydrogen combines with the oxygen of air.

Because it combines with oxygen, hydrogen is capable of removing oxygen from certain substances. For example, when hydrogen is passed over hot cupric oxide, the latter is converted to reddish-brown copper.

$$\underset{\text{Cupric oxide}}{CuO} + \underset{\text{Hydrogen}}{H_2} \longrightarrow \underset{\text{Copper}}{Cu} + \underset{\text{Water}}{H_2O}$$

This type of reaction is called reduction. Hydrogen, which removes oxygen, is called a reducing agent.

Uses

A mixture of hydrogen and oxygen burns giving out much heat. This is used in the oxy-hydrogen flame which has a high temperature (about 2400°C) and melts most metals.

Hydrogen is used for filling weather balloons. Hydrogen is used to convert vegetable oils like groundnut and cotton-seed into solid fats (Vanaspati). This process is called hydrogenation of oils

Hydrogen also finds use in the manufacture of ammonia.

6.7 Common salt

Apart from air and water, another widely distributed substance is common salt or sodium chloride. It is an essential component of our food. It is the chief constituent of sea-water in which it occurs to an extent of about 3 percent by weight. It also occurs as rock salt, a mineral. Our blood also contains salt.

Common salt is used in industries such as ice-making, pottery, paper making, caustic soda, washing soda, hydrochloric acid, etc. It is also used as a preservative.

6.8 Chlorine

Chlorine is an important constituent of common salt. It is used for the sterilization of water and as a bleaching agent in paper and textile industries.

Preparation

Chlorine can be prepared by the action of an oxidising agent like manganese dioxide or potassium permanganate on concentrated hydrochloric acid.

$$\underset{\text{Manganese dioxide}}{MnO_2} + \underset{\text{Hydrochloric acid}}{4\,HCl} \longrightarrow \underset{\text{Manganous chloride}}{MnCl_2} + \underset{\text{Water}}{2\,H_2O} + \underset{\text{Chlorine}}{Cl_2\uparrow}$$

A mixture of sodium chloride and concentrated sulphuric acid may be used to get hydrochloric acid.

$NaCl$	+	$H_2SO_4 \longrightarrow$	$NaHSO_4$	+	HCl
Sodium chloride		Sulphuric acid	Sodium bisulphate		Hydrochloric acid

In the laboratory, therefore, chlorine is prepared by heating a mixture of sodium chloride and manganese dioxide with concentrated sulphuric acid. The greenish yellow gas evolved is collected by the upward displacement of air (Fig. 6.9).

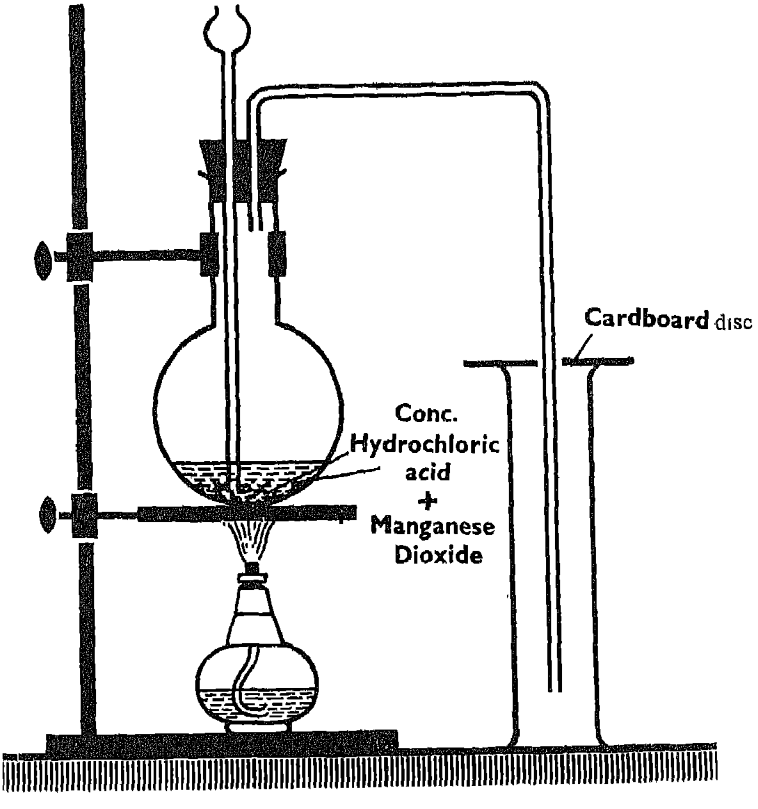


Figure 6.9 *Preparation of chlorine*

Properties

Chlorine is a greenish yellow gas with an irritating smell. It affects the nose, throat and lungs. It is poisonous. It is fairly soluble in water. Since chlorine can be collected by upward displacement of air, it can be inferred that chlorine is heavier than air.

Chlorine reacts vigorously with a number of substances. For example, when a jar of chlorine is inverted over a jar of hydrogen in diffuse light, the greenish yellow colour of chlorine disappears. The product formed (hydrogen chloride) gives dense white fumes of ammonium chloride with ammonia.

$$\begin{array}{ccccc} H_2 & + & Cl_2 & \longrightarrow & 2HCl \\ \text{Hydrogen} & & \text{Chlorine} & & \text{Hydrogen chloride} \end{array}$$

$$\begin{array}{ccccc} HCl & + & NH_3 & \longrightarrow & NH_4Cl \\ \text{Hydrogen} & & \text{Ammonia} & & \text{Ammonium chloride} \\ \text{chloride} & & & & \end{array}$$

An aqueous solution of hydrogen chloride gas is called hydrochloric acid.

Sodium burns in chlorine forming sodium chloride.

$$\begin{array}{ccccc} 2Na & + & Cl_2 & \longrightarrow & 2NaCl \\ \text{Sodium} & & \text{Chlorine} & & \text{Sodium chloride} \end{array}$$

Similarly burning magnesium continues to burn in chlorine, forming magnesium chloride.

$$\begin{array}{ccccc} Mg & + & Cl_2 & \longrightarrow & MgCl_2 \\ \text{Magnesium} & & \text{Chlorine} & & \text{Magnesium chloride} \end{array}$$

Heated iron filings catch fire in chlorine, forming ferric chloride.

$$\begin{array}{ccccc} 2Fe & + & 3Cl_2 & \longrightarrow & 2FeCl_3 \\ \text{Iron} & & \text{Chlorine} & & \text{Ferric chloride} \end{array}$$

A burning candle (which contains carbon and hydrogen), when lowered into a jar of chlorine, continues to burn, depositing carbon particles

Other elements like antimony, copper, etc , also react vigorously with chlorine forming respective chlorides. Thus chlorine is a very reactive element.

Chlorine in the presence of moisture, bleaches litmus and other coloured matter.

Uses

Chlorine is used for bleaching cotton, linen, paper, pulp, etc. It is used to kill harmful bacteria and to disinfect town water supply. It is used in the manufacture of a number of useful compounds like bleaching powder (a disinfectant), chloroform (an anaesthetic), carbon tetrachloride (a solvent and a fire extinguisher) and D. D. T. (an insecticide).

6.9 Carbon

All living things, both plants and animals, contain compounds of carbon. Many materials of daily use like petrol, coal and wood also contain carbon. The element carbon exists in various forms like diamond, graphite, charcoal, soot etc.

Forms of carbon

Diamond which has an octahedral crystalline form exhibits great brilliance after cutting and polishing. Because of this property diamond is used in making jewels. Diamond is extremely hard and so is used in cutting glass sheets and precious stones and in drilling rocks.

Graphite, another form of carbon, is black, soft and flaky. It is used as a lubricant for machinery. The 'lead' of pencils is graphite mixed with clay. Graphite is a good conductor of electricity.

Charcoal is obtained by burning wood in a limited supply of air. It is a black, porous solid. It is insolube in water. A piece of charcoal, when dropped into water, floats. But when boiled in water, it sinks. This is because, on heating, the air present in the pores is driven out, causing the charcoal to sink.

When powdered charcoal is added to a gas jar containing reddish-brown vapours of nitrogen dioxide, the colour disappears after sometime. The colour of litmus solution also disappears on boiling with charcoal powder. Charcoal thus takes up nitrogen dioxide vapour and the litmus. A special type of charcoal called activated charcoal does this better. This property is made use of in gasmasks.

When charcoal burns in an adequate supply of air or oxygen, carbon dioxide is produced.

$$\underset{\text{Carbon}}{C} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2\uparrow}$$

When a mixture of powdered charcoal and cupric oxide is heated strongly, carbon dioxide and metallic copper are formed. In this reaction charcoal acts as a reducing agent and cupric oxide as an oxidising agent

$$\underset{\text{Cupric oxide}}{2CuO} + \underset{\text{Carbon}}{C} \longrightarrow \underset{\text{Copper}}{2Cu} + \underset{\text{Carbon dioxide}}{CO_2\uparrow}$$

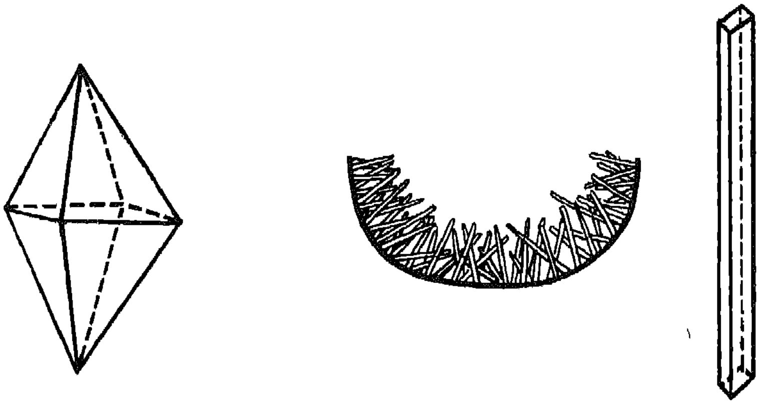
6.10 Sulphur

Sulphur is widely distributed in nature. Deposits of free sulphur occur in volcanic regions. In the combined state, it occurs as sulphides, e. g., lead sulphide, zinc sulphide and mercuric sulphide. It also occurs in the form of sulphates, e. g., gypsum (calcium sulphate) and epsom salt (magnesium sulphate).

Forms of sulphur

Like carbon, sulphur also exists in different forms — rhombic, prismatic and plastic. The first two are crystalline.

The familiar form of sulphur is rhombic. It is stable at room temperature. It has a bright yellow colour. It is insoluble in water but soluble in carbon disulphide and carbon tetrachloride. When heated to a temperature below 100°C it changes to prismatic sulphur.



Rhomtic

Prismatic

Figure 6.10 *Forms of sulphur*

Prismatic sulphur is dull yellow. It is also insoluble in water but soluble both in carbon disulphide and carbon tetrachloride.

Sulphur on heating melts to a brown liquid. When the boiling liquid is poured into cold water, a soft rubbery mass of plastic sulphur is obtained. Unlike crystalline sulphur, plastic sulphur is insoluble in both carbon disulphide and carbon tetrachloride.

All the forms of sulphur burn in air, forming sulphur dioxide, a gas with a suffocating smell.

$$\underset{\text{Sulphur}}{S} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Sulphur dioxide}}{SO_2\uparrow}$$

Sulphur reacts with metals like copper and iron when heated forming their sulphides.

$$\underset{\text{Copper}}{Cu} + \underset{\text{Sulphur}}{S} \longrightarrow \underset{\text{Cupric sulphide}}{CuS}$$

$$\underset{\text{Iron}}{Fe} + \underset{\text{Sulphur}}{S} \longrightarrow \underset{\text{Ferrous sulphide}}{FeS}$$

Uses

Sulphur is mainly used in the manufacture of sulphuric acid, one of the most important chemicals.

Natural rubber is soft and sticky. To make it hard and strong it is heated with sulphur. This process is known as vulcanization of rubber. Sulphur is used in making gun-powder, fire works and matches. Sulphur ointment and sulpha drugs are used in the treatment of certain diseases. Sulphur also finds use as a fungicide, insecticide and fumigant.

Questions

1. What gases are removed from air by
   a. burning of wood b. respiration c. photosynthesis?
2. How would you be affected if the percentage of oxygen and nitrogen in air were just reversed?
3. If a fish is placed in water which has been previously boiled and cooled, it will soon die. Explain.

4. How would you detect the presence of carbon dioxide and water-vapour in air?
5. What is the role of manganese dioxide in the preparation of oxygen from potassium chlorate?
6. How would you obtain oxygen and hydrogen from water?
7. In the electrolysis of water, what are the relative volumes of the gases evolved?
8. Give a test for oxygen.
9. For laboratory preparation of oxygen and nitrogen give the following ·
   a. materials used
   b. equation for the reaction taking place and
   c. the method of collection.
10. What is the product formed when hydrogen is burnt in air?
11. Give a test for hydrogen.
12. What is reduction? Illustrate with examples.
13. Name three important properties of hydrogen.
14. Is water a compound or a mixture?
15. Name two substances which can burn in chlorine
16. Which crystalline form of carbon is
   a. very hard b soft and greasy?
17. What form of carbon is used for decolourising brown sugar?
18. How would you test whether a given substance is sulphur?
19. Answer the following :
   a. If a fire breaks out, which of the following gases would you use to fight the fire?
      i) hydrogen ii) chlorine iii) carbon dioxide iv) oxygen.
   b. A room is filled with oxygen. What will happen when a match stick is lighted in such a room?
   c. How would you know that oxygen is dissolved in water?
   d. Jars containing hydrogen, chlorine, carbon dioxide and nitrogen are covered with lids. If the lids are removed, which of the gases would escape most easily? How would you identify each gas?
   e. In jars containing different gases, a burning candle is slowly lowered. If the observation is as follows, infer whether the gas is hydrogen, oxygen, carbon dioxide or chlorine.
      i) The flame is put out.
      ii) The flame is put out but the gas burns.
      iii) The flame becomes brighter.
      iv) The candle continues to burn depositing carbon.

f. A gas is collected by upward displacement of air. What property of the gas does this indicate?

g. A gas is collected over water. What property of the gas does this indicate?

20. Fill up the blanks with suitable words.

a. Oxygen is obtained in the laboratory by heating . . . .

b. Nitrogen is obtained when a mixture of. . .. . . and .. . . is heated.

c. Carbon dioxide turns ... . milky.

d. ...'... ... . is the lightest gas known.

e When hydrogen burns in air,. ... . .is formed.

f. In the reaction of hydrogen with copper oxide to form copper and water, hydrogen is called a.. .... agent.

g When potassium permanganate reacts with concentrated hydrochloric acid, the gas liberated is.... . ...

h. Chlorine is .... than air.

i. When hydrogen and chlorine react.. . ...is formed.

j. Chlorine . ... colouring matter in the presence of...... ....

k. .. .. is the hardest of all substances.

l. Graphite is a good conductor of ....

m. In some chemical reactions, carbon acts as a.. .. ..agent.

n. When sulphur burns in air ... . . is formed.

o. Sulphur is insoluble in . . .. but soluble in.... . ...

21. Choose the right answer.

a. After phosphorus is burnt in a closed bell-jar kept over water, the water level rises approximately to

i) 1/2 ii) 3/4 iii) 1/5 the original volume.

b. The gas in the atmosphere most essential for photosynthesis is

i) oxygen ii) carbon dioxide iii) nitrogen.

c. The following gas is ordinarily found free in air

i) chlorine ii) hydrogen iii) nitrogen

d Manganese dioxide, when heated with potassium chlorate

i) remains unchanged ii) is completely changed iii) is partially changed after the reaction

e. A molecule of oxygen is represented by

i) O ii) O_2 iii) 2O

f. The gas most soluble in water under normal conditions is

i) carbon dioxide ii) oxygen iii) hydrogen

g. Chlorine at room temperature is

i) a gas ii) a liquid iii) a solid

h. The following gas has an irritating smell

i) oxygen ii) carbon dioxide iii) chlorine

i. The following is a bleaching agent

i) oxygen ii) chlorine iii) hydrogen

j. The following is a reducing agent

i) oxygen ii) carbon dioxide iii) hydrogen

k. A test for hydrogen is that

i) it bleaches moist coloured flowers
ii) it rekindles a glowing splinter
iii) it puts out a burning splinter with a 'pop' sound and burns with a pale blue flame.

l. The following gives a very hot flame on burning with oxygen

i) hydrogen ii) carbon dioxide iii) nitrogen

22. Write against each item in group II the appropriate matching item from group I.

GROUP I	GROUP II
a. Fire extinguisher	Charcoal
b. Hydrogenation	Diamond
c. Lubricant	Chlorine
d. Hardest substance	Hydrogen
e. Sterilization of water	Graphite
f. Not a reactive gas	Carbon dioxide
	Nitrogen

GROUP I	GROUP II
a. Graphite	Catalyst in the preparation of oxygen
b. Sulphur	Turns lime water milky
c. Carbon dioxide	Hardest substance
d. Hydrogen	Soluble in carbon disulphide
c. Marble	Calcium carbonate
f. Manganese dioxide	Occurs in air
	Soft and greasy
	Lightest substance.

23. State whether the following statements are TRUE or FALSE.

a. Air is a compound, not a mixture.
b. Oxygen forms 1/6 of air by volume.
c. Oxygen is insoluble in water.
d. Nitrogen is not a reactive gas.
e. Carbon dioxide turns lime water milky.
f. Carbon dioxide enters air by photosynthesis.
g. Burning magnesium is put out by nitrogen.
h. Chlorine is a bleaching agent.
i. Charcoal floats on water
j. Sulphur is soluble in water.

UNIT 7

Common Metals and Alloys

Metals and alloys are widely used in everyday life. Metals possess characteristic properties and the specific uses to which they are put to largely depend on these properties. In this unit the reactivity of metals towards air and water, their occurrence in nature and one of the methods used for obtaining them from their ores, are dealt with. Some of the common alloys and their uses are also discussed.

In every home we find vessels of aluminium, brass, copper and stainless steel. Copper and aluminium wires are used in making electric cables. Electric bulbs contain a filament of tungsten metal. Automobiles, trains and aeroplanes require large quantities of different metals in their construction. Thus we find that metals are widely used in everyday life. Modern facilities of transport and communication would not have been possible without them.

7.1 How do we recognise metals?

Metals in pure state appear bright. This property is called metallic lustre. It is this property that lends beauty to objects made of metals. Silver is white, gold yellow and copper reddish-brown. Metals like copper and aluminium when exposed to air acquire a dull appearance. This is due to the formation of an oxide, carbonate or sulphide film on the surface of the metal. These films are removed by rubbing with sand paper. Tamarind and lemon can also remove these films. That is why they are used in cleaning metalware.

Most of the metals are solids at room temperature. Mercury, which is a liquid, is an exception. Metals are usually heavy. However, a few metals like sodium and potassium are lighter than water.

Metals are generally hard. The hardness varies from metal to metal. Sodium is soft enough to be cut with a knife whereas iron is very hard.

Some metals can be rolled or beaten into thin sheets. This property is called malleability. Gold is the most malleable metal. It is possible to get gold sheets so thin that 100,000 of these placed one over the other make only about a centimeter thick pile

Some metals are very ductile, that is, they can be drawn into very fine wires For example, a wire of about 2 kilometer length can be obtained from only one gram of gold.

Malleability and ductility of metals enable us to fashion them into a variety of shapes required for different purposes. For example, thin foils of silver and gold are used to cover sweets and aluminium foil to wrap chocolates, seal milk bottles and pack medicinal tablets.

Metals conduct heat. It is observed that different metals have different capacities for conducting heat. Metals like copper and

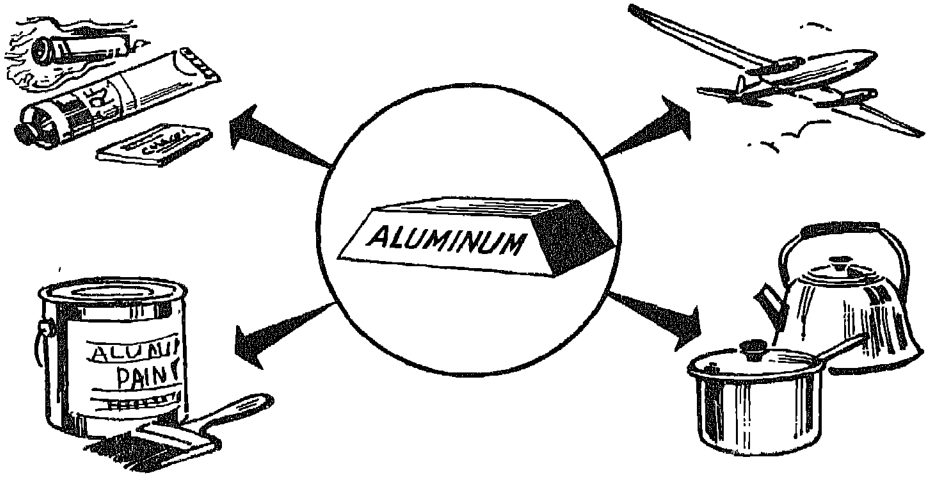


Figure 7.1 *Aluminium and its uses*

aluminium being good conductors of heat are used in making cooking vessels, water-boilers and in other heating or cooling systems. Metals conduct electricity also. Copper and aluminium, being good conductors of electricity, are used for transmitting electric power.

7.2 Reactivity of metals

We learnt in unit 3 (page 39) that an atom of sodium loses an electron to form a sodium ion (Na^{+}). Similarly a magnesium atom loses two electrons to form a magnesium ion (Mg^{++}). In general metals have a tendency to lose one or more electrons to form positive ions.

a. Action of air

A freshly cut piece of sodium has a shining bluish white appearance However, it soon becomes dull on exposure to air. This is due to the formation of an oxide film on the surface by reaction with air. Magnesium and zinc also tarnish in air but rather slowly. Bright aluminium and copper retain their lustre for longer periods. Gold remains unaffected in air. Thus metals show varying reactivity towards air.

We learnt in unit 5 that reactions in general go faster at higher temperatures than at lower temperatures. Hence when metals are heated in air, they react much faster forming oxides.

Molten sodium burns with a yellow flame and forms sodium monoxide and sodium peroxide.

$$\underset{\text{Sodium}}{4\,Na} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Sodium monoxide}}{2\,Na_2O}$$

$$\underset{\text{Sodium}}{2\,Na} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Sodium peroxide}}{Na_2O_2}$$

Magnesium burns with a dazzling white light giving magnesium oxide.

$$\underset{\text{Magnesium}}{2\,Mg} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Magnesium oxide}}{2\,MgO}$$

Aluminium, copper and iron do not burn in air on heating. However, on heating, they burn readily in oxygen forming oxides.

$$\underset{\text{Aluminium}}{4\,Al} + \underset{\text{Oxygen}}{3\,O_2} \longrightarrow \underset{\text{Aluminium oxide}}{2\,Al_2O_3}$$

$$\underset{\text{Copper}}{2\,Cu} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Cupric oxide}}{2\,CuO}$$

$$\underset{\text{Iron}}{3\,Fe} + \underset{\text{Oxygen}}{2\,O_2} \longrightarrow \underset{\text{Magnetic oxide of iron}}{Fe_3O_4(FeO+Fe_2O_3)}$$

Thus we find that many metals which do not burn in air on heating, burn readily in oxygen. Some metals like gold, however, remain unaffected even when heated in oxygen Such metals are called noble metals.

b. Action of water

Sodium reacts vigorously with cold water forming hydrogen and sodium hydroxide.

$$\underset{\text{Sodium}}{2\,Na} + \underset{\text{Water}}{2\,H_2O} \longrightarrow \underset{\text{Sodium hydroxide}}{2\,NaOH} + \underset{\text{Hydrogen}}{H_2\uparrow}$$

Magnesium, on the other hand, reacts slowly with cold water. However, it reacts readily with boiling water liberating hydrogen.

$$\underset{\text{Magnesium}}{Mg} + \underset{\text{Water}}{2\,H_2O} \longrightarrow \underset{\text{Magnesium hydroxide}}{Mg\,(OH)_2} + \underset{\text{Hydrogen}}{H_2\uparrow}$$

Copper, silver and gold are not acted upon by boiling water or even steam. Metals thus differ in their reactivity towards water.

c. Corrosion of metals

Many metals tarnish or acquire a coating on their surfaces on exposure to moist air. For instance, iron is covered with a brown coating. In a similar manner, copper acquires a green coating. This phenomenon is known as corrosion of metals Losses due to corrosion of metals amount to millions of rupees each year.

Corrosion can be prevented if contact between air and the metal is cut off. There are several ways of doing this. The most common method is to coat the metal surface with a paint. That is why we find that steel furniture, railway over-bridges, bodies of motor cars, railway coaches, etc., are always painted. Sometimes coating is done with molten metals. Thus iron sheets used in roofing are dipped into molten zinc after careful cleaning. Such sheets are known as galvanised iron sheets 'Tin cans' are made of thin iron sheets coated with tin. Cooking vessels of brass or copper acquire a green coating in moist air or in contact with certain food materials This green coating is poisonous. Cooking vessels are, therefore, 'tinned' (coated with tin).

Metals like aluminium and chromium form oxide films which adhere to the metal and thus protect the underlying metal from further corrosion.

7.3 How do metals occur in nature?

As seen earlier, sodium is very reactive. It combines readily with other elements to form compounds. Hence it does not occur free in nature. This is true of all active metals. Sodium is found mainly in the form of sodium chloride (common salt). Noble metals like gold, however, occur in the free state.

In the combined state, metals generally occur as oxides, sulphides or carbonates mixed with varying amounts of rocky material. Such naturally occurring materials are known as ores. For example, iron and aluminium occur as oxides (ferric oxide, Fe_2O_3 and aluminium oxide, Al_2O_3); copper, zinc and lead occur as sulphides (copper iron sulphide, $CuFeS_2$, zinc sulphide, ZnS and lead sulphide, PbS).

7·4 Obtaining a metal from its ore

Cerussite (lead carbonate), an ore of lead, when heated decomposes into lead oxide and carbon dioxide

$$\underset{\text{Lead carbonate}}{PbCO_3} \longrightarrow \underset{\text{Lead oxide}}{PbO} \longrightarrow + \underset{\text{Carbon dioxide}}{CO_2\uparrow}$$

Lead oxide so formed, is reduced to lead on heating with charcoal which acts as a reducing agent.

$$\underset{\text{Lead oxide}}{PbO} + \underset{\text{Carbon}}{C} \longrightarrow \underset{\text{Lead}}{Pb} + \underset{\text{Carbon monoxide}}{CO\uparrow}$$

7.5 Alloys

A pure metal may not possess enough mechanical strength or resistance to corrosion to meet the demands of everyday use. For example, pure gold is soft and wears out easily For making ornaments gold is made hard by adding copper or silver. Thus pure metals are often mixed with other elements. Stainless steel contains mostly iron and small quantities of chromium and nickel. Brass contains copper and zinc. Materials such as stainless steel and brass are called **alloys.** Alloys ordinarily contain two or more metals. The components and uses of a few common alloys are given below.

Table 7.1 Some common alloys

Name	Components	Uses
Brass	Cu and Zn	Utensils, ornaments
Bronze	Cu, Sn and Zn	Statues
Solder	Pb and Sn	Soldering metals
Stainless steel	Fe, Cr and Ni	Utensils, machinery
Magnalium	Mg and Al	Aircraft, balance beams

Alloys may be prepared by heating together molten metals in the required proportion until a homogeneous mixture is obtained.

Questions

1. Give an example of
   a. a metal which is liquid at ordinary temperature,
   b. a metal which does not tarnish in air,
   c. a metal which is lighter than water.
2. Why does gold occur free in nature?
3. Mention four metals (or alloys) used in daily life. What are their uses?
4. Arrange the following metals in order of their reactivity with water.
   Gold, sodium, magnesium
5. Fill up the blanks.
   a. Metals can be beaten into thin sheets. This property is called... .. . ..
   b Metals can be drawn into wires. This property is called......... ..
   c. Active metals occur in . .. . .state in nature.
   d. Lead is obtained by. .. . . of lead oxide.
   e. Brass is an alloy of ... .

6. Tick off the correct answers in the following.
   - i) Gold is called a noble metal because
     - a. it has an attractive yellow colour
     - b. it is hard and heavy
     - c. it is unreactive
     - d. it is used in making ornaments.
   - ii) Only those metals occur free in nature which
     - a. are hard
     - b. occur in large amounts
     - c. are reactive
     - d. are unreactive.
   - iii) The best conductor of heat is
     - a. iron
     - b. aluminium
     - c copper

7. Of the metals sodium, magnesium, aluminium which
   - a. reacts vigorously with cold water?
   - b reacts only with boiling water?
   - c. occurs in the free state in nature?
   - d. is the softest metal?

8. State whether the following statements are TRUE or FALSE.
   - a. All metals are silvery white.
   - b. Sodium reacts with water forming sodium hydroxide.
   - c. Iron is a better conductor of heat than copper.
   - d. Copper is alloyed with gold to make gold softer.
   - e. Brass is an alloy of copper and zinc.
   - f. Magnalium is an alloy of zinc and lead.
   - g. Magnesium burns in air forming magnesium oxide.
   - h. Noble metals occur free in nature.

9. A solid substance is soft and conducts electricity.

   It can be cut with a knife. The freshly cut piece is bright but it tarnishes immediately.

   The substance is
   - a. copper
   - b. silver
   - c. aluminium
   - d. sodium
   - e. gold

10. Give reasons why

a. steel furniture is generally painted.

b. cooking vessels made of brass are tinned.

c. copper is usually added to gold while making ornaments.

d. aluminium is used in making electric cables,

e. sodium does not occur free in nature.

11. Complete the following equations:

a. $Na + H_2O \longrightarrow NaOH + \ldots\ldots\ldots$

b $Mg + \ldots\ldots\ldots \longrightarrow MgSO_4 + H_2$

c. $Al + O_2 \longrightarrow \ldots\ldots\ldots$

d $PbCO_3 \longrightarrow \ldots\ldots\ldots + CO_2$

e. $PbO + C \longrightarrow \ldots\ldots\ldots + \ldots\ldots\ldots$

UNIT 8

Acids, Bases and Salts

Acids, bases and salts play an important role in chemical reactions. They are extensively used in laboratory and industry. In this unit we shall learn about their formation and properties.

8.1 Acids and bases

We come across substances such as vinegar, lemon juice, tamarind and orange juice, which have a sour taste We are also familiar with the slippery feel of solutions of soap and washing soda. The substances with sour taste turn blue litmus paper red. Such substances are called **acids.** The substances with slippery feel turn red litmus paper blue and turmeric paper reddish-brown. These are called **bases.**

Acids and bases are widely used in laboratory and industry. Some acids and bases which are commonly employed in the laboratory are given below.

Table 8.1 Some common acids and bases

Acid	Formula	Base	Formula
Acetic acid	CH_3COOH	Ammonium hydroxide	NH_4OH
Hydrochloric acid	HCl	Calcium hydroxide	$Ca(OH)_2$
Nitric acid	HNO_3	Magnesium hydroxide	$Mg(OH)_2$
Phosphoric acid	H_3PO_4	Potassium hydroxide	KOH
Sulphuric acid	H_2SO_4	Sodium hydroxide	$NaOH$

8.2 Formation of acids and bases

When carbon, sulphur and phosphorus are burnt in air they form oxides. These oxides react with water forming solutions which turn blue litmus paper red showing that they are acidic. Most *acids are, therefore, formed by the reaction of oxides of elements such as carbon, sulphur and phosphorus with water.*

When magnesium and sodium are burnt in air, they also form oxides. These oxides react with water forming solutions or suspensions which turn red litmus paper blue showing that they are basic. *Bases are, therefore, formed when oxides of metals react with water.*

8.3 Properties of acids and bases

Aqueous solutions of acids and bases have the following properties.

Acids	Bases
1. Have a sour taste	Have a bitter taste
2. Turn blue litmus paper red	Turn red litmus paper blue
3. React with bases forming salt and water	React with acids forming salt and water
4. Conduct electricity	Conduct electricity
5. Liberate hydrogen when metals such as magnesium, zinc and iron react with them	Have soapy touch
6. Decompose carbonates, evolving carbon dioxide	Turn turmeric paper reddish-brown

Acids give hydrogen ions, $\overset{+}{H}$, whereas bases give hydroxyl ions, $\overset{-}{OH}$, in aqueous solutions. Hydrogen and hydroxyl ions are responsible for the similarity in some of the properties of acids and bases respectively.

Bases like sodium hydroxide and potassium hydroxide which are highly soluble in water are called **alkalies.**

8.4 Neutralization

It is observed that when an *appropriate* amount of sodium hydroxide is added to a given quantity of hydrochloric acid, the resulting solution is *neutral*, that is, it is neither acidic nor basic. The solution, on evaporation, gives a solid salt, sodium chloride.

$$\mathrm{NaOH} + \mathrm{HCl} \longrightarrow \mathrm{NaCl} + \mathrm{H_2O}$$

Such a reaction of a base with an acid is called **neutralization.** Thus acids and bases react forming salt and water.

Sometimes we experience a burning sensation in the stomach. This is due to the formation of excess hydrochloric acid in the stomach. To neutralize the excess acid, mild bases like aluminium hydroxide, $Al(OH)_3$, or milk of magnesia — magnesium hydroxide, $Mg(OH)_2$, are used.

When soils become much acidic, they lose fertility. Slaked lime or limestone is generally added to neutralize the acid and restore the fertility of the soil.

A bee-sting causes a burning sensation. This is due to the injection of an acid into the body. To relieve the burning sensation, the acid should be neutralized. This is done by rubbing moist baking soda or ammonia water on the spot where the bee has stung.

8.5 Salts

We have learnt that a salt is formed when an acid is neutralized by a base. Common salt (sodium chloride) is a product of mutual neutralization of hydrochloric acid and sodium hydroxide Some salts formed from different acids and bases are given below.

Table 8.2 Some important salts

Bases Acids	$NaOH$	KOH	NH_4OH	$Ca(OH)_2$
HCl	$NaCl$	KCl	NH_4Cl	$CaCl_2$
HNO_3	$NaNO_3$	KNO_3	NH_4NO_3	$Ca(NO_3)_2$
H_2SO_4	Na_2SO_4	K_2SO_4	$(NH_4)_2SO_4$	$CaSO_4$
H_2CO_3	Na_2CO_3	K_2CO_3	$(NH_4)_2CO_3$	$CaCO_3$

Salts are widely used in daily life. A few salts with their uses are listed below.

Table 8.3 Some salts and their uses

a. Used in household

Salt	Formula	Use
Sodium bicarbonate	$NaHCO_3$	Baking
Sodium carbonate	Na_2CO_3	Washing
Sodium chloride	$NaCl$	To give taste to food

b. Used in medicine

Barium sulphate	$BaSO_4$	X-ray photography
Calcium sulphate	$CaSO_4$	Casts, moulds, plasters
Ferrous sulphate	$FeSO_4$	Tonic
Magnesium sulphate	$MgSO_4$	Purgative
Magnesium carbonate	$MgCO_3$	Antacid
Potassium bromide	KBr	Sedative
Potassium permanganate	$KMnO_4$	Antiseptic

Questions

1. Vinegar and lemon juice taste sour. What other properties do you expect them to have in common?
2. Name three common acids used in the laboratory. Write their formulae. What is common to all of them?
3. How are acids formed? Give their important properties.
4. What are the distinguishing properties of acids and alkalies?
5. Write the names and formulae of four common bases. What do they have in common?
6. Mention two practical applications of neutralization.

7. If you were given sodium hydroxide, ammonium hydroxide, nitric acid and hydrochloric acid, how many salts can you prepare by neutralization? Give their names.

8. Write equations for the action of water on

   a. CaO

   b. CO_2

   c. MgO

   Name the product in each case.

9. Explain why

   a. a stain of turmeric turns reddish-brown when soap is applied

   b. sodium carbonate solution is used to wash the floor on which acid is spilled.

10. Mention some of the salts which you come across in daily life.

11. If you were given three bottles labelled A, B and C and were told that one contained hydrochloric acid, another sodium hydroxide solution and the third water, how would you identify them?

12. By mistake a student got some sodium hydroxide solution in his mouth After washing his mouth with a lot of water, he rinsed it with vinegar and then finally with water Explain why he did this.

13. State whether the following statements are TRUE or FALSE.

    a. Oxides of metals react with water forming acids.

    b. Magnesium reacts with acids and gives hydrogen.

    c. Aqueous solutions of bases decompose carbonates evolving carbon dioxide.

    d. Aqueous solutions of acids conduct electricity.

    e. Passing carbon dioxide into lime water is a process of neutralization.

15. Which of the following properties would a solution have if it is an acid?

    a. Solution turns red litmus paper blue.

    b. Solution tastes sour.

    c. Solution has a slippery feel.

16. Which of the following properties are shown by both acids and bases?

    a. Solution turns litmus red

    b Solution tastes sour.

    c. Solution is a conductor of electricity.

    d. Solution reacts with magnesium to liberate hydrogen

17. When appropriate amounts of an acid and a base are mixed

    a. the properties of the acid alone are lost.

    b. the properties of the base alone are lost.

    c. the properties of both the acid and the base are lost.

    d. the properties of neither the acid nor the base are lost

18. Substance 'A' has been obtained by reacting potassium hydroxide with hydrochloric acid. Its solution is neutral to litmus. What is the substance 'A'? Give its formula.

19. Fill up the blanks

a. Phosphorus burns in air forming . .. . . . which dissolves in water to give a solution which turns .. .. litmus paper ....

b. Potassium burns in air forming .... which reacts with water to give a solution which turns . ... litmus paper .. . .

c. The reaction of an acid with a base is called . .. .

d. Sodium chloride may be obtained by treating ... . with hydrochloric acid.

20. Match the following :

A	B
Has a slippery feel	Vinegar
Reaction between an acid and a base	Barium sulphate
Turns blue litmus paper red	Alkali
Used in X-ray photography	Water
Does not change the colour of litmus	Sodium chloride

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